

# CHAPTER 5 GASES AND THE KINETIC-MOLECULAR THEORY

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## END-OF-CHAPTER PROBLEMS

- 5.1 Plan: Review the behavior of the gas phase vs. the liquid phase.  
Solution:  
a) The volume of the liquid remains constant, but the volume of the gas increases to the volume of the larger container.  
b) The volume of the container holding the gas sample increases when heated, but the volume of the container holding the liquid sample remains essentially constant when heated.  
c) The volume of the liquid remains essentially constant, but the volume of the gas is reduced.
- 5.2 The particles in a gas are further apart than those are in a liquid.  
a) The greater empty space between gas molecules allows gases to be more compressible than liquids.  
b) The greater empty space between gas molecules allows gases to flow with less resistance (hindrance) than liquids.  
c) The large empty space between gas molecules limits their interaction, allowing all mixtures of gases to be solutions.  
d) The large empty space between gas molecules increases the volume of the gas, therefore decreasing the density.
- 5.3 The mercury column in the mercury barometer stays up due to the force exerted by the atmosphere on the mercury in the outer reservoir just balancing the gravitational force on the mercury in the tube. Its height adjusts according to the air pressure on the reservoir. The column of mercury is shorter on a mountaintop as there is less atmosphere to exert a force on the mercury reservoir. On a mountaintop, the air pressure is less, so the height of mercury it balances in the barometer is shorter than at sea level where there is more air pressure.
- 5.4 Plan: The ratio of the heights of columns of mercury and water are inversely proportional to the ratio of the densities of the two liquids. Convert the height in mm to height in cm.  
Solution:  
$$\frac{h_{\text{H}_2\text{O}}}{h_{\text{Hg}}} = \frac{d_{\text{Hg}}}{d_{\text{H}_2\text{O}}}$$
$$h_{\text{H}_2\text{O}} = \frac{d_{\text{Hg}}}{d_{\text{H}_2\text{O}}} \times h_{\text{Hg}} = \left( \frac{13.5 \text{ g/mL}}{1.00 \text{ g/mL}} \right) (730 \text{ mmHg}) \left( \frac{10^{-3} \text{ m}}{1 \text{ mm}} \right) \left( \frac{1 \text{ cm}}{10^{-2} \text{ m}} \right) = 985.5 = \mathbf{990 \text{ cm H}_2\text{O}}$$
- 5.5 Plan: The ratio of the heights of columns of mercury and water are inversely proportional to the ratio of the densities of the two liquids.  
Solution:  
$$\frac{h_{\text{H}_2\text{O}}}{h_{\text{Hg}}} = \frac{d_{\text{Hg}}}{d_{\text{H}_2\text{O}}}$$
$$h_{\text{H}_2\text{O}} = \frac{d_{\text{Hg}}}{d_{\text{H}_2\text{O}}} \times h_{\text{Hg}} = \left( \frac{13.5 \text{ g/mL}}{1.00 \text{ g/mL}} \right) (755 \text{ mmHg}) = 10,192.5 = \mathbf{1.02 \times 10^4 \text{ mm H}_2\text{O}}$$
- 5.6 Plan: Use the conversion factors between pressure units:  
1 atm = 760 mmHg = 760 torr = 101.325 kPa = 1.01325 bar  
Solution:  
a) Converting from atm to mmHg:  $P(\text{mmHg}) = (0.745 \text{ atm}) \left( \frac{760 \text{ mmHg}}{1 \text{ atm}} \right) = 566.2 = \mathbf{566 \text{ mmHg}}$

- b) Converting from torr to bar:  $P(\text{bar}) = (992 \text{ torr}) \left( \frac{1.01325 \text{ bar}}{760 \text{ torr}} \right) = 1.32256 = \mathbf{1.32 \text{ bar}}$
- c) Converting from kPa to atm:  $P(\text{atm}) = (365 \text{ kPa}) \left( \frac{1 \text{ atm}}{101.325 \text{ kPa}} \right) = 3.60227 = \mathbf{3.60 \text{ atm}}$
- d) Converting from mmHg to kPa:  $P(\text{kPa}) = (804 \text{ mmHg}) \left( \frac{101.325 \text{ kPa}}{760 \text{ mmHg}} \right) = 107.191 = \mathbf{107 \text{ kPa}}$

5.7 Plan: Use the conversion factors between pressure units:

$$1 \text{ atm} = 760 \text{ mmHg} = 760 \text{ torr} = 101.325 \text{ kPa} = 1.01325 \text{ bar}$$

Solution:

a) Converting from cmHg to atm:

$$P(\text{atm}) = (76.8 \text{ cmHg}) \left( \frac{10^{-2} \text{ m}}{1 \text{ cm}} \right) \left( \frac{1 \text{ mm}}{10^{-3} \text{ m}} \right) \left( \frac{1 \text{ atm}}{760 \text{ mmHg}} \right) = 1.01053 = \mathbf{1.01 \text{ atm}}$$

b) Converting from atm to kPa:  $P(\text{kPa}) = (27.5 \text{ atm}) \left( \frac{101.325 \text{ kPa}}{1 \text{ atm}} \right) = 2.786 \times 10^3 = \mathbf{2.79 \times 10^3 \text{ kPa}}$

c) Converting from atm to bar:  $P(\text{bar}) = (6.50 \text{ atm}) \left( \frac{1.01325 \text{ bar}}{1 \text{ atm}} \right) = 6.5861 = \mathbf{6.59 \text{ bar}}$

d) Converting from kPa to torr:  $P(\text{torr}) = (0.937 \text{ kPa}) \left( \frac{760 \text{ torr}}{101.325 \text{ kPa}} \right) = 7.02808 = \mathbf{7.03 \text{ torr}}$

5.8 Plan: Use the conversion factors between pressure units:

$$1 \text{ atm} = 760 \text{ mmHg} = 760 \text{ torr} = 1.01325 \times 10^5 \text{ Pa} = 14.7 \text{ psi}$$

Solution:

a) Converting from mmHg to atm:  $P(\text{atm}) = (2.75 \times 10^2 \text{ mmHg}) \left( \frac{1 \text{ atm}}{760 \text{ mmHg}} \right) = 0.361842 = \mathbf{0.362 \text{ atm}}$

b) Converting from psi to atm:  $P(\text{atm}) = (86 \text{ psi}) \left( \frac{1 \text{ atm}}{14.7 \text{ psi}} \right) = 5.85034 = \mathbf{5.9 \text{ atm}}$

c) Converting from Pa to atm:  $P(\text{atm}) = (9.15 \times 10^6 \text{ Pa}) \left( \frac{1 \text{ atm}}{1.01325 \times 10^5 \text{ Pa}} \right) = 90.303 = \mathbf{90.3 \text{ atm}}$

d) Converting from torr to atm:  $P(\text{atm}) = (2.54 \times 10^4 \text{ torr}) \left( \frac{1 \text{ atm}}{760 \text{ torr}} \right) = 33.42105 = \mathbf{33.4 \text{ atm}}$

5.9 Plan:  $1 \text{ atm} = 1.01325 \times 10^5 \text{ Pa} = 1.01325 \times 10^5 \text{ N/m}^2$ . So the force on  $1 \text{ m}^2$  of ocean is  $1.01325 \times 10^5 \text{ N}$  where

$$1 \text{ N} = 1 \frac{\text{kg} \cdot \text{m}}{\text{s}^2}. \text{ Use } F = mg \text{ to find the mass of the atmosphere in kg/m}^2 \text{ for part a). For part b), convert this mass}$$

to  $\text{g/cm}^2$  and use the density of osmium to find the height of this mass of osmium.

Solution:

a)  $F = mg$

$$1.01325 \times 10^5 \text{ N} = mg$$

$$1.01325 \times 10^5 \frac{\text{kg} \cdot \text{m}}{\text{s}^2} = (\text{mass}) (9.81 \text{ m/s}^2)$$

$$\text{mass} = 1.03287 \times 10^4 = \mathbf{1.03 \times 10^4 \text{ kg}}$$

b)  $\left( 1.03287 \times 10^4 \frac{\text{kg}}{\text{m}^2} \right) \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left( \frac{10^{-2} \text{ m}}{1 \text{ cm}} \right)^2 = 1.03287 \times 10^3 \text{ g/cm}^2$

$$\text{Height (cm)} = \left( 1.03287 \times 10^3 \frac{\text{g}}{\text{cm}^2} \right) \left( \frac{1 \text{ mL}}{22.6 \text{ g}} \right) \left( \frac{1 \text{ cm}^3}{1 \text{ mL}} \right) = 45.702 = \mathbf{45.7 \text{ cm Os}}$$

5.10 The statement is incomplete with respect to temperature and mass of sample. The correct statement is: At constant temperature and moles of gas, the volume of gas is inversely proportional to the pressure.

- 5.11 a) Charles's law: At constant pressure, the volume of a fixed amount of gas is directly proportional to its Kelvin temperature. **Variable: volume and temperature; Fixed: pressure and moles**  
 b) Avogadro's law: At fixed temperature and pressure, the volume occupied by a gas is directly proportional to the moles of gas. **Variable: volume and moles; Fixed: temperature and pressure**  
 c) Amontons's law: At constant volume, the pressure exerted by a fixed amount of gas is directly proportional to the Kelvin temperature. **Variable: pressure and temperature; Fixed: volume and moles**

5.12 Plan: Examine the ideal gas law; volume and temperature are constant and pressure and moles are variable.  
Solution:

$$PV = nRT \quad P = n \frac{RT}{V} \quad R, T, \text{ and } V \text{ are constant}$$

$$P = n \times \text{constant}$$

At constant temperature and volume, the pressure of the gas is directly proportional to the amount of gas in moles.

5.13 Plan: Examine the ideal gas law, noting the fixed variables and those variables that change.  $R$  is always constant  
 so  $\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$ .

Solution:

a)  $P$  is fixed; both  $V$  and  $T$  double:  $\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$  or  $\frac{V_1}{n_1 T_1} = \frac{V_2}{n_2 T_2}$

$T$  can double as  $V$  doubles only if  $n$  is **fixed**.

b)  $T$  and  $n$  are both fixed and  $V$  doubles:  $\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$  or  $P_1 V_1 = P_2 V_2$

$P$  and  $V$  are inversely proportional; as  $V$  doubles,  $P$  is **halved**.

c)  $T$  is fixed and  $V$  doubles.  $n$  doubles since one mole of reactant gas produces a total of 2 moles of product gas.

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2} \quad \text{or} \quad \frac{P_1 V_1}{n_1} = \frac{P_2 V_2}{n_2}$$

$V$  and  $n$  can both double only if  $P$  is **fixed**.

d)  $P$  is fixed and  $V$  doubles.  $n$  is fixed since 2 moles of reactant gas produce 2 moles of product gas.

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2} \quad \text{or} \quad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$V$  and  $T$  are directly proportional so as  $V$  is doubled,  $T$  is **doubled**.

5.14 Plan: Use the relationship  $\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$  or  $V_2 = \frac{P_1 V_1 n_2 T_2}{P_2 n_1 T_1}$ .

Solution:

a) As the pressure on a fixed amount of gas ( $n$  is fixed) increases at constant temperature ( $T$  is fixed), the molecules move closer together, decreasing the volume. When the pressure is tripled, the **volume decreases to one-third of the original volume** at constant temperature (Boyle's law).

$$V_2 = \frac{P_1 V_1 n_2 T_2}{P_2 n_1 T_1} = \frac{(P_1)(V_1)(1)(1)}{(3P_1)(1)(1)} \quad V_2 = \frac{1}{3} V_1$$

b) As the temperature of a fixed amount of gas ( $n$  is fixed) increases at constant pressure ( $P$  is fixed), the gas molecules gain kinetic energy. With higher energy, the gas molecules collide with the walls of the container with greater force, which increases the size (volume) of the container. If the temperature is increased by a factor of 3.0

(at constant pressure) then the **volume will increase by a factor of 3.0** (Charles's law).

$$V_2 = \frac{P_1 V_1 n_2 T_2}{P_2 n_1 T_1} = \frac{(1)(V_1)(1)(3T_1)}{(1)(1)(T_1)} \quad V_2 = 3V_1$$

c) As the number of molecules of gas increases at constant pressure and temperature ( $P$  and  $T$  are fixed), the force they exert on the container increases. This results in an increase in the volume of the container. Adding 3 moles of gas to 1 mole increases the number of moles by a factor of 4, thus the **volume increases by a factor of 4** (Avogadro's law).

$$V_2 = \frac{P_1 V_1 n_2 T_2}{P_2 n_1 T_1} = \frac{(1)(V_1)(4n_1)(1)}{(1)(n_1)(1)} \quad V_2 = 4V_1$$

- 5.15 Plan: Use the relationship  $\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$  or  $V_2 = \frac{P_1 V_1 T_2}{P_2 T_1}$ .  $R$  and  $n$  are fixed.

Solution:

a) As the pressure on a fixed amount of gas ( $n$  is fixed) decreases at constant temperature ( $T$  is fixed), the molecules move farther together, increasing the volume. When the pressure is reduced by a factor of 4, the **volume increases by a factor of 4** at constant temperature (Boyle's law).

$$V_2 = \frac{P_1 V_1 T_2}{P_2 T_1} = \frac{(P_1)(V_1)(1)}{(1/4P_1)(1)} \quad V_2 = 4V_1$$

b) As the pressure on a fixed amount of gas ( $n$  is fixed) doubles from 101 kPa to 202 kPa at constant temperature, the volume decreases by a factor of  $\frac{1}{2}$ . As the temperature of a fixed amount of gas ( $n$  is fixed) decreases by a factor of  $\frac{1}{2}$  (from 310 K to 155 K) at constant pressure, the volume decreases by a factor of  $\frac{1}{2}$ . The changes in pressure and temperature combine to **decrease the volume by a factor of 4**.

$$P_1 = 760 \text{ torr} = 101 \text{ kPa} \quad T_1 = 37^\circ\text{C} + 273 = 310 \text{ K}$$

$$V_2 = \frac{P_1 V_1 T_2}{P_2 T_1} = \frac{(101 \text{ kPa})(V_1)(155 \text{ K})}{(202 \text{ kPa})(310 \text{ K})} \quad V_2 = \frac{1}{4} V_1$$

c) As the pressure on a fixed amount of gas ( $n$  is fixed) decreases at constant temperature ( $T$  is fixed), the molecules move farther together, increasing the volume. When the pressure is reduced by a factor of 2, the **volume increases by a factor of 2** at constant temperature (Boyle's law).

$$T_2 = 32^\circ\text{C} + 273 = 305 \text{ K} \quad P_2 = 101 \text{ kPa} = 1 \text{ atm}$$

$$V_2 = \frac{P_1 V_1 T_2}{P_2 T_1} = \frac{(2 \text{ atm})(V_1)(305 \text{ K})}{(1 \text{ atm})(305 \text{ K})} \quad V_2 = 2V_1$$

- 5.16 Plan: This is Charles's law: at constant pressure and with a fixed amount of gas, the volume of a gas is directly proportional to the absolute temperature of the gas. The temperature must be lowered to reduce the volume of a gas. Arrange the ideal gas law, solving for  $T_2$  at fixed  $n$  and  $P$ . Temperature must be converted to kelvin.

Solution:

$$V_1 = 9.10 \text{ L} \quad V_2 = 2.50 \text{ L}$$

$$T_1 = 198^\circ\text{C} \text{ (convert to K)} \quad T_2 = \text{unknown}$$

$n$  and  $P$  remain constant

$$\text{Converting } T \text{ from } ^\circ\text{C to K: } T_1 = 198^\circ\text{C} + 273 = 471 \text{ K}$$

Arranging the ideal gas law and solving for  $T_2$ :

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2} \quad \text{or} \quad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$T_2 = T_1 \frac{V_2}{V_1} = 471 \text{ K} \left( \frac{2.50 \text{ L}}{9.10 \text{ L}} \right) = 129.396 \text{ K} - 273 = -143.604 = \mathbf{-144^\circ\text{C}}$$

- 5.17 Plan: This is Charles's law: at constant pressure and with a fixed amount of gas, the volume of a gas is directly proportional to the absolute temperature of the gas. If temperature is reduced, the volume of gas will also be reduced. Arrange the ideal gas law, solving for  $V_2$  at fixed  $n$  and  $P$ . Temperature must be converted to kelvins.

Solution:

$$V_1 = 93 \text{ L}$$

$$V_2 = \text{unknown}$$

$$T_1 = 145^\circ\text{C} \text{ (convert to K)}$$

$$T_2 = -22^\circ\text{C}$$

$n$  and  $P$  remain constant

$$\text{Converting } T \text{ from } ^\circ\text{C to K: } T_1 = 145^\circ\text{C} + 273 = 418 \text{ K}$$

$$T_2 = -22^\circ\text{C} + 273 = 251 \text{ K}$$

Arranging the ideal gas law and solving for  $V_2$ :

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2} \quad \text{or} \quad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$V_2 = V_1 \frac{T_2}{T_1} = 93 \text{ L} \left( \frac{251 \text{ K}}{418 \text{ K}} \right) = 55.844 = \mathbf{56 \text{ L}}$$

- 5.18 Plan: Since the volume, temperature, and pressure of the gas are changing, use the combined gas law. Arrange the ideal gas law, solving for  $V_2$  at fixed  $n$ . STP is  $0^\circ\text{C}$  (273 K) and 1 atm (101.325 kPa)

Solution:

$$P_1 = 153.3 \text{ kPa}$$

$$P_2 = 101.325 \text{ kPa}$$

$$V_1 = 25.5 \text{ L}$$

$$V_2 = \text{unknown}$$

$$T_1 = 298 \text{ K}$$

$$T_2 = 273 \text{ K}$$

$n$  remains constant

Arranging the ideal gas law and solving for  $V_2$ :

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2} \quad \text{or} \quad \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$V_2 = V_1 \left( \frac{T_2}{T_1} \right) \left( \frac{P_1}{P_2} \right) = (25.5 \text{ L}) \left( \frac{273 \text{ K}}{298 \text{ K}} \right) \left( \frac{153.3 \text{ kPa}}{101.325 \text{ kPa}} \right) = 35.3437 = \mathbf{35.3 \text{ L}}$$

- 5.19 Plan: Since the volume, temperature, and pressure of the gas are changing, use the combined gas law. Arrange the ideal gas law, solving for  $V_2$  at fixed  $n$ . Temperature must be converted to kelvins.

Solution:

$$P_1 = 745 \text{ torr}$$

$$P_2 = 367 \text{ torr}$$

$$V_1 = 3.65 \text{ L}$$

$$V_2 = \text{unknown}$$

$$T_1 = 298 \text{ K}$$

$$T_2 = -14^\circ\text{C} + 273 = 259 \text{ K}$$

$n$  remains constant

Arranging the ideal gas law and solving for  $V_2$ :

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2} \quad \text{or} \quad \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$V_2 = V_1 \left( \frac{T_2}{T_1} \right) \left( \frac{P_1}{P_2} \right) = (3.65 \text{ L}) \left( \frac{259 \text{ K}}{298 \text{ K}} \right) \left( \frac{745 \text{ torr}}{367 \text{ torr}} \right) = 6.4397 = \mathbf{6.44 \text{ L}}$$

- 5.20 Plan: Given the volume, pressure, and temperature of a gas, the number of moles of the gas can be calculated using the ideal gas law, solving for  $n$ . The gas constant,  $R = 0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$ , gives pressure in atmospheres and temperature in Kelvin. The given pressure in torr must be converted to atmospheres and the temperature converted to kelvins.

Solution:

$$P = 328 \text{ torr (convert to atm)}$$

$$V = 5.0 \text{ L}$$

$$T = 37^\circ\text{C}$$

$$n = \text{unknown}$$

$$\text{Converting } P \text{ from torr to atm: } P = (328 \text{ torr}) \left( \frac{1 \text{ atm}}{760 \text{ torr}} \right) = 0.43158 \text{ atm}$$

$$\text{Converting } T \text{ from } ^\circ\text{C to K:}$$

$$T = 37^\circ\text{C} + 273 = 310 \text{ K}$$

$$PV = nRT$$

Solving for  $n$ :

$$n = \frac{PV}{RT} = \frac{(0.43158 \text{ atm})(5.0 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(310 \text{ K})} = 0.08479 = \mathbf{0.085 \text{ mol chlorine}}$$

- 5.21 Plan: Given the volume, moles, and temperature of a gas, the pressure of the gas can be calculated using the ideal gas law, solving for  $P$ . The gas constant,  $R = 0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$ , gives volume in liters and temperature in Kelvin. The given volume in mL must be converted to L and the temperature converted to kelvins.

Solution:

$$V = 75.0 \text{ mL}$$

$$T = 26^\circ\text{C}$$

$$n = 1.47 \times 10^{-3} \text{ mol}$$

$$P = \text{unknown}$$

$$\text{Converting } V \text{ from mL to L: } V = (75.0 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right) = 0.0750 \text{ L}$$

$$\text{Converting } T \text{ from } ^\circ\text{C} \text{ to K: } T = 26^\circ\text{C} + 273 = 299 \text{ K}$$

$$PV = nRT$$

Solving for  $P$ :

$$P = \frac{nRT}{V} = \frac{(1.47 \times 10^{-3} \text{ mol})\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(299 \text{ K})}{0.0750 \text{ L}} = 0.48114 \text{ atm}$$

$$\text{Convert } P \text{ to units of torr: } (0.48114 \text{ atm})\left(\frac{760 \text{ torr}}{1 \text{ atm}}\right) = 365.6664 = \mathbf{366 \text{ torr}}$$

- 5.22 Plan: Solve the ideal gas law for moles and convert to mass using the molar mass of  $\text{ClF}_3$ . The gas constant,  $R = 0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$ , gives volume in liters, pressure in atmospheres, and temperature in Kelvin so volume must be converted to L, pressure to atm, and temperature to K.

Solution:

$$V = 357 \text{ mL}$$

$$T = 45^\circ\text{C}$$

$$P = 699 \text{ mmHg}$$

$$n = \text{unknown}$$

$$\text{Converting } V \text{ from mL to L: } V = (357 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right) = 0.357 \text{ L}$$

$$\text{Converting } T \text{ from } ^\circ\text{C} \text{ to K: } T = 45^\circ\text{C} + 273 = 318 \text{ K}$$

$$\text{Converting } P \text{ from mmHg to atm: } P = (699 \text{ mmHg})\left(\frac{1 \text{ atm}}{760 \text{ mmHg}}\right) = 0.91974 \text{ atm}$$

$$PV = nRT$$

Solving for  $n$ :

$$n = \frac{PV}{RT} = \frac{(0.91974 \text{ atm})(0.357 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(318 \text{ K})} = 0.01258 \text{ mol ClF}_3$$

$$\text{Mass ClF}_3 = (0.01258 \text{ mol ClF}_3) \left(\frac{92.45 \text{ g ClF}_3}{1 \text{ mol ClF}_3}\right) = 1.163021 = \mathbf{1.16 \text{ g ClF}_3}$$

- 5.23 Plan: Solve the ideal gas law for pressure; convert mass to moles using the molar mass of  $\text{N}_2\text{O}$ . The gas constant,  $R = 0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$ , gives temperature in Kelvin so the temperature must be converted to units of kelvins.

Solution:

$$V = 3.1 \text{ L}$$

$$T = 115^\circ\text{C}$$

$$n = 75.0 \text{ g (convert to moles)}$$

$$P = \text{unknown}$$

$$\text{Converting } T \text{ from } ^\circ\text{C} \text{ to K:}$$

$$T = 115^\circ\text{C} + 273 = 388 \text{ K}$$

Converting from mass of  $\text{N}_2\text{O}$  to moles:  $n = (75.0 \text{ g N}_2\text{O}) \left( \frac{1 \text{ mol N}_2\text{O}}{44.02 \text{ g N}_2\text{O}} \right) = 1.70377 \text{ mol N}_2\text{O}$

$$PV = nRT$$

Solving for  $P$ :

$$P = \frac{nRT}{V} = \frac{(1.70377 \text{ mol}) \left( 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (388 \text{ K})}{(3.1 \text{ L})} = 17.5075 = \mathbf{18 \text{ atm N}_2\text{O}}$$

- 5.24 Plan: Solve the ideal gas law for moles. The gas constant,  $R = 0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$ , gives pressure in atmospheres, and temperature in Kelvin so pressure must be converted to atm and temperature to K.

Solution:

$$V = 1.5 \text{ L}$$

$$T = 23^\circ\text{C}$$

$$P = 85 + 14.7 = 99.7 \text{ psi}$$

$$n = \text{unknown}$$

Converting  $T$  from  $^\circ\text{C}$  to K:

$$T = 23^\circ\text{C} + 273 = 296 \text{ K}$$

$$\text{Converting } P \text{ from psi to atm: } P = (99.7 \text{ psi}) \left( \frac{1 \text{ atm}}{14.7 \text{ psi}} \right) = 6.7823 \text{ atm}$$

$$PV = nRT$$

Solving for  $n$ :

$$n = \frac{PV}{RT} = \frac{(6.7823 \text{ atm})(1.5 \text{ L})}{\left( 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (296 \text{ K})} = 0.41863 = \mathbf{0.42 \text{ mol SO}_2}$$

- 5.25 Air is mostly  $\text{N}_2$  (28.02 g/mol),  $\text{O}_2$  (32.00 g/mol), and argon (39.95 g/mol). These “heavy” gases dominate the density of dry air. Moist air contains  $\text{H}_2\text{O}$  (18.02 g/mol). The relatively light water molecules lower the density of the moist air.

- 5.26 The molar mass of  $\text{H}_2$  is less than the average molar mass of air (mostly  $\text{N}_2$ ,  $\text{O}_2$ , and Ar), so air is denser. To collect a beaker of  $\text{H}_2(g)$ , **invert** the beaker so that the air will be replaced by the lighter  $\text{H}_2$ . The molar mass of  $\text{CO}_2$  is greater than the average molar mass of air, so  $\text{CO}_2(g)$  is more dense. Collect the  $\text{CO}_2$  holding the beaker **upright**, so the lighter air will be displaced out the top of the beaker.

- 5.27 Gases mix to form a solution and each gas in the solution behaves as if it were the only gas present.

- 5.28  $P_A = X_A P_T$  The partial pressure of a gas ( $P_A$ ) in a mixture is directly proportional to its mole fraction ( $X_A$ ).

- 5.29 Plan: Calculate the mole fraction of each gas; the partial pressure of each gas is directly proportional to its mole fraction so the gas with the highest mole fraction has the highest partial pressure. Use the relationship between partial pressure and mole fraction to calculate the partial pressure of gas  $\text{D}_2$ .

Solution:

$$\text{a) } X_A = \frac{n_A}{n_{\text{total}}} = \frac{4 \text{ A particles}}{16 \text{ total particles}} = 0.25$$

$$X_B = \frac{n_B}{n_{\text{total}}} = \frac{3 \text{ B particles}}{16 \text{ total particles}} = 0.1875$$

$$X_C = \frac{n_C}{n_{\text{total}}} = \frac{5 \text{ C particles}}{16 \text{ total particles}} = 0.3125$$

$$X_{\text{D}_2} = \frac{n_{\text{D}_2}}{n_{\text{total}}} = \frac{4 \text{ D}_2 \text{ particles}}{16 \text{ total particles}} = 0.25$$

**Gas C** has the highest mole fraction and thus the highest partial pressure.

b) **Gas B** has the lowest mole fraction and thus the lowest partial pressure.

$$\text{c) } P_{\text{D}_2} = X_{\text{D}_2} \times P_{\text{total}} \quad P_{\text{D}_2} = 0.25 \times 0.75 \text{ atm} = 0.1875 = \mathbf{0.19 \text{ atm}}$$

- 5.30 Plan: Rearrange the ideal gas law to calculate the density of xenon from its molar mass at STP. Standard temperature is  $0^\circ\text{C}$  (273 K) and standard pressure is 1 atm. Do not forget that the pressure at STP is exact and will not affect the significant figures.

Solution:

$$P = 1 \text{ atm}$$

$$T = 273 \text{ K}$$

$$\mathcal{M} \text{ of Xe} = 131.3 \text{ g/mol}$$

$$d = \text{unknown}$$

$$PV = nRT$$

$$\text{Rearranging to solve for density: } d = \frac{P\mathcal{M}}{RT} = \frac{(1 \text{ atm})(131.3 \text{ g/mol})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(273 \text{ K})} = 5.8581 = \mathbf{5.86 \text{ g/L}}$$

- 5.31 Plan: Rearrange the ideal gas law to calculate the density of  $\text{CFCl}_3$  from its molar mass. Temperature must be converted to kelvins.

Solution:

$$P = 1.5 \text{ atm}$$

$$T = 120^\circ\text{C} + 273 = 393 \text{ K}$$

$$\mathcal{M} \text{ of CFCl}_3 = 137.4 \text{ g/mol}$$

$$d = \text{unknown}$$

$$PV = nRT$$

Rearranging to solve for density:

$$d = \frac{P\mathcal{M}}{RT} = \frac{(1.5 \text{ atm})(137.4 \text{ g/mol})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(393 \text{ K})} = 6.385807663 = \mathbf{6.4 \text{ g/L}}$$

- 5.32 Plan: Solve the ideal gas law for moles. Convert moles to mass using the molar mass of  $\text{AsH}_3$  and divide this mass by the volume to obtain density in g/L. Standard temperature is  $0^\circ\text{C}$  (273 K) and standard pressure is 1 atm. Do not forget that the pressure at STP is exact and will not affect the significant figures.

Solution:

$$V = 0.0400 \text{ L}$$

$$T = 0^\circ\text{C} + 273 = 273 \text{ K}$$

$$P = 1 \text{ atm}$$

$$n = \text{unknown}$$

$$\mathcal{M} \text{ of AsH}_3 = 77.94 \text{ g/mol}$$

$$PV = nRT$$

Solving for  $n$ :

$$n = \frac{PV}{RT} = \frac{(1 \text{ atm})(0.0400 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(273 \text{ K})} = 1.78465 \times 10^{-3} = \mathbf{1.78 \times 10^{-3} \text{ mol AsH}_3}$$

Converting moles of  $\text{AsH}_3$  to mass of  $\text{AsH}_3$ :

$$\text{Mass (g) of AsH}_3 = (1.78465 \times 10^{-3} \text{ mol AsH}_3) \left( \frac{77.94 \text{ g AsH}_3}{1 \text{ mol AsH}_3} \right) = 0.1391 \text{ g AsH}_3$$

$$d = \frac{\text{mass}}{\text{volume}} = \frac{(0.1391 \text{ g})}{(0.0400 \text{ L})} = 3.4775 = \mathbf{3.48 \text{ g/L}}$$

- 5.33 Plan: Solve the density form of the ideal gas law for molar mass. Temperature must be converted to kelvins. Compare the calculated molar mass to the molar mass values of the noble gases to determine the identity of the gas.

Solution:

$$P = 3.00 \text{ atm}$$

$$T = 0^\circ\text{C} + 273 = 273 \text{ K}$$

$$d = 2.71 \text{ g/L}$$

$$\mathcal{M} = \text{unknown}$$

$$d = \frac{P\mathcal{M}}{RT}$$

Rearranging to solve for molar mass:

$$\mathcal{M} = \frac{dRT}{P} = \frac{(2.71 \text{ g/L}) \left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right) (273 \text{ K})}{(3.00 \text{ atm})} = 20.24668 = \mathbf{20.2 \text{ g/mol}}$$

Therefore, the gas is **Ne**.

- 5.34 Plan: Rearrange the formula  $PV = (m/M)RT$  to solve for molar mass. Convert the mass in ng to g and volume in  $\mu\text{L}$  to L. Temperature must be in Kelvin and pressure in atm.

Solution:

$$V = 0.206 \mu\text{L}$$

$$T = 45^\circ\text{C} + 273 = 318 \text{ K}$$

$$P = 380 \text{ torr}$$

$$m = 206 \text{ ng}$$

$M$  = unknown

$$\text{Converting } P \text{ from torr to atm: } P = (380 \text{ torr}) \left( \frac{1 \text{ atm}}{760 \text{ torr}} \right) = 0.510526 \text{ atm}$$

$$\text{Converting } V \text{ from } \mu\text{L} \text{ to L: } V = (0.206 \mu\text{L}) \left( \frac{10^{-6} \text{ L}}{1 \mu\text{L}} \right) = 2.06 \times 10^{-7} \text{ L}$$

$$\text{Converting } m \text{ from ng to g: } m = (206 \text{ ng}) \left( \frac{10^{-9} \text{ g}}{1 \text{ ng}} \right) = 2.06 \times 10^{-7} \text{ g}$$

$$PV = \left( \frac{m}{M} \right) RT$$

Solving for molar mass,  $M$ :

$$M = \frac{mRT}{PV} = \frac{(2.06 \times 10^{-7} \text{ g}) \left( 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (318 \text{ K})}{(0.510526 \text{ atm}) (2.06 \times 10^{-7} \text{ L})} = 51.1390 = \mathbf{51.1 \text{ g/mol}}$$

- 5.35 Plan: Rearrange the formula  $PV = (m/M)RT$  to solve for molar mass. Compare the calculated molar mass to that of  $\text{N}_2$ , Ne, and Ar to determine the identity of the gas. Convert volume to liters, pressure to atm, and temperature to Kelvin.

Solution:

$$V = 63.8 \text{ mL}$$

$$T = 22^\circ\text{C} + 273 = 295 \text{ K}$$

$$P = 747 \text{ mm Hg}$$

$$m = 0.103 \text{ g}$$

$M$  = unknown

$$\text{Converting } P \text{ from mmHg to atm: } P = (747 \text{ mmHg}) \left( \frac{1 \text{ atm}}{760 \text{ mmHg}} \right) = 0.982895 \text{ atm}$$

$$\text{Converting } V \text{ from mL to L: } V = (63.8 \text{ mL}) \left( \frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) = 0.0638 \text{ L}$$

$$PV = \left( \frac{m}{M} \right) RT$$

Solving for molar mass,  $M$ :

$$M = \frac{mRT}{PV} = \frac{(0.103 \text{ g}) \left( 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (295 \text{ K})}{(0.982895 \text{ atm}) (0.0638 \text{ L})} = 39.7809 = 39.8 \text{ g/mol}$$

The molar masses are  $\text{N}_2 = 28 \text{ g/mol}$ , Ne = 20 g/mol, and Ar = 40 g/mol. Therefore, the gas is **Ar**.

- 5.36 Plan: Use the ideal gas law to determine the number of moles of Ar and of  $\text{O}_2$ . The gases are combined ( $n_{\text{total}} = n_{\text{Ar}} + n_{\text{O}_2}$ ) into a 400 mL flask ( $V$ ) at  $27^\circ\text{C}$  ( $T$ ). Use the ideal gas law again to determine the total pressure from  $n_{\text{total}}$ ,  $V$ , and  $T$ . Pressure must be in units of atm, volume in units of L and temperature in K.

Solution:

For Ar:

$$V = 0.600 \text{ L}$$

$$T = 227^\circ\text{C} + 273 = 500. \text{ K}$$

$$P = 1.20 \text{ atm}$$

$$n = \text{unknown}$$

$$PV = nRT$$

Solving for  $n$ :

$$n = \frac{PV}{RT} = \frac{(1.20 \text{ atm})(0.600 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(500. \text{ K})} = 0.017539586 \text{ mol Ar}$$

For  $\text{O}_2$ :

$$V = 0.200 \text{ L}$$

$$T = 127^\circ\text{C} + 273 = 400. \text{ K}$$

$$P = 501 \text{ torr}$$

$$n = \text{unknown}$$

Converting  $P$  from torr to atm:  $P = (501 \text{ torr})\left(\frac{1 \text{ atm}}{760 \text{ torr}}\right) = 0.6592105 \text{ atm}$

$$PV = nRT$$

Solving for  $n$ :

$$n = \frac{PV}{RT} = \frac{(0.6592105 \text{ atm})(0.200 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(400. \text{ K})} = 0.004014680 \text{ mol O}_2$$

$$n_{\text{total}} = n_{\text{Ar}} + n_{\text{O}_2} = 0.017539586 \text{ mol} + 0.004014680 \text{ mol} = 0.021554266 \text{ mol}$$

For the mixture of Ar and  $\text{O}_2$ :

$$V = 400 \text{ mL}$$

$$T = 27^\circ\text{C} + 273 = 300. \text{ K}$$

$$P = \text{unknown}$$

$$n = 0.021554265 \text{ mol}$$

Converting  $V$  from mL to L:  $V = (400 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right) = 0.400 \text{ L}$

$$PV = nRT$$

Solving for  $P$ :

$$P_{\text{mixture}} = \frac{nRT}{V} = \frac{(0.021554266 \text{ mol})\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(300 \text{ K})}{(0.400 \text{ L})} = 1.32720 = \mathbf{1.33 \text{ atm}}$$

- 5.37 **Plan:** Use the ideal gas law, solving for  $n$  to find the total moles of gas. Convert the mass of Ne to moles and subtract moles of Ne from the total number of moles to find moles of Ar. Volume must be in units of liters, pressure in units of atm, and temperature in kelvins.

**Solution:**

$$V = 355 \text{ mL}$$

$$T = 35^\circ\text{C} + 273 = 308 \text{ K}$$

$$P = 626 \text{ mmHg}$$

$$n_{\text{total}} = \text{unknown}$$

Converting  $P$  from mmHg to atm:  $P = (626 \text{ mmHg})\left(\frac{1 \text{ atm}}{760 \text{ mmHg}}\right) = 0.823684 \text{ atm}$

Converting  $V$  from mL to L:  $V = (355 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right) = 0.355 \text{ L}$

$$PV = nRT$$

Solving for  $n_{\text{total}}$ :

$$n_{\text{total}} = \frac{PV}{RT} = \frac{(0.823684 \text{ atm})(0.355 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(308 \text{ K})} = 0.011563655 \text{ mol Ne} + \text{mol Ar}$$

$$\text{Moles Ne} = (0.146 \text{ g Ne})\left(\frac{1 \text{ mol Ne}}{20.18 \text{ g Ne}}\right) = 0.007234886 \text{ mol Ne}$$

$$\text{Moles Ar} = n_{\text{total}} - n_{\text{Ne}} = (0.011563655 - 0.007234886) \text{ mol} = 0.004328769 = \mathbf{0.0043 \text{ mol Ar}}$$

- 5.38 Plan: Use the ideal gas law, solving for  $n$  to find the moles of  $O_2$ . Use the molar ratio from the balanced equation to determine the moles (and then mass) of phosphorus that will react with the oxygen. Standard temperature is  $0^\circ\text{C}$  (273 K) and standard pressure is 1 atm.

Solution:

$$V = 35.5 \text{ L}$$

$$T = 0^\circ\text{C} + 273 = 273 \text{ K}$$

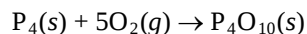
$$P = 1 \text{ atm}$$

$$n = \text{unknown}$$

$$PV = nRT$$

Solving for  $n$ :

$$n = \frac{PV}{RT} = \frac{(1 \text{ atm})(35.5 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(273 \text{ K})} = 1.583881 \text{ mol } O_2$$



$$\text{Mass } P_4 = (1.583881 \text{ mol } O_2) \left( \frac{1 \text{ mol } P_4}{5 \text{ mol } O_2} \right) \left( \frac{123.88 \text{ g } P_4}{1 \text{ mol } P_4} \right) = 39.24224 = \mathbf{39.2 \text{ g } P_4}$$

- 5.39 Plan: Use the ideal gas law, solving for  $n$  to find the moles of  $O_2$  produced. Volume must be in units of liters, pressure in atm, and temperature in kelvins. Use the molar ratio from the balanced equation to determine the moles (and then mass) of potassium chlorate that reacts.

Solution:

$$V = 638 \text{ mL}$$

$$T = 128^\circ\text{C} + 273 = 401 \text{ K}$$

$$P = 752 \text{ torr}$$

$$n = \text{unknown}$$

$$\text{Converting } P \text{ from torr to atm: } P = (752 \text{ torr}) \left( \frac{1 \text{ atm}}{760 \text{ torr}} \right) = 0.9894737 \text{ atm}$$

$$\text{Converting } V \text{ from mL to L: } V = (638 \text{ mL}) \left( \frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) = 0.638 \text{ L}$$

$$PV = nRT$$

Solving for  $n$ :

$$n = \frac{PV}{RT} = \frac{(0.9894737 \text{ atm})(0.638 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(401 \text{ K})} = 0.0191751 \text{ mol } O_2$$



$$\text{Mass (g) of } KClO_3 = (0.0191751 \text{ mol } O_2) \left( \frac{2 \text{ mol } KClO_3}{3 \text{ mol } O_2} \right) \left( \frac{122.55 \text{ g } KClO_3}{1 \text{ mol } KClO_3} \right) = 1.5666 = \mathbf{1.57 \text{ g } KClO_3}$$

- 5.40 Plan: Since the amounts of two reactants are given, this is a limiting reactant problem. To find the mass of  $PH_3$ , write the balanced equation and use molar ratios to find the number of moles of  $PH_3$  produced by each reactant. The smaller number of moles of product indicates the limiting reagent. Solve for moles of  $H_2$  using the ideal gas law.

Solution:

Moles of hydrogen:

$$V = 83.0 \text{ L}$$

$$T = 0^\circ\text{C} + 273 = 273 \text{ K}$$

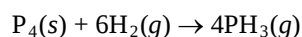
$$P = 1 \text{ atm}$$

$$n = \text{unknown}$$

$$PV = nRT$$

Solving for  $n$ :

$$n = \frac{PV}{RT} = \frac{(1 \text{ atm})(83.0 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(273 \text{ K})} = 3.7031584 \text{ mol } H_2$$



$$\text{PH}_3 \text{ from P}_4 = (37.5 \text{ g P}_4) \left( \frac{1 \text{ mol P}_4}{123.88 \text{ g P}_4} \right) \left( \frac{4 \text{ mol PH}_3}{1 \text{ mol P}_4} \right) = 1.21085 \text{ mol PH}_3$$

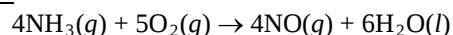
$$\text{PH}_3 \text{ from H}_2 = (3.7031584 \text{ mol H}_2) \left( \frac{4 \text{ mol PH}_3}{6 \text{ mol H}_2} \right) = 2.4687723 \text{ mol PH}_3$$

P<sub>4</sub> is the limiting reactant because it forms less PH<sub>3</sub>.

$$\text{Mass PH}_3 = (37.5 \text{ g P}_4) \left( \frac{1 \text{ mol P}_4}{123.88 \text{ g P}_4} \right) \left( \frac{4 \text{ mol PH}_3}{1 \text{ mol P}_4} \right) \left( \frac{33.99 \text{ g PH}_3}{1 \text{ mol PH}_3} \right) = 41.15676 = \mathbf{41.2 \text{ g PH}_3}$$

- 5.41 Plan: Since the amounts of two reactants are given, this is a limiting reactant problem. To find the mass of NO, write the balanced equation and use molar ratios to find the number of moles of NO produced by each reactant. Since the moles of gas are directly proportional to the volumes of the gases at the same temperature and pressure, the limiting reactant may be found by comparing the volumes of the gases. The smaller volume of product indicates the limiting reagent. Then use the ideal gas law to convert the volume of NO produced to moles and then to mass.

Solution:



$$\text{Mol NO from NH}_3 = (35.6 \text{ L NH}_3) \left( \frac{4 \text{ L NO}}{4 \text{ L NH}_3} \right) = 35.6 \text{ L NO}$$

$$\text{Mol NO from O}_2 = (40.5 \text{ L O}_2) \left( \frac{4 \text{ L NO}}{5 \text{ L O}_2} \right) = 32.4 \text{ L NO}$$

O<sub>2</sub> is the limiting reactant since it forms less NO.

Converting volume of NO to moles and then mass:

$$V = 32.4 \text{ L}$$

$$T = 0^\circ\text{C} + 273 = 273 \text{ K}$$

$$P = 1 \text{ atm}$$

$$n = \text{unknown}$$

$$PV = nRT$$

Solving for  $n$ :

$$n = \frac{PV}{RT} = \frac{(1 \text{ atm})(32.4 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(273 \text{ K})} = 1.44557 \text{ mol NO}$$

$$\text{Mass (g) of NO} = (1.44557 \text{ mol NO}) \left( \frac{30.01 \text{ g NO}}{1 \text{ mol NO}} \right) = 43.38156 = \mathbf{43.4 \text{ g NO}}$$

- 5.42 Plan: First, write the balanced equation. The moles of hydrogen produced can be calculated from the ideal gas law. The problem specifies that the hydrogen gas is collected over water, so the partial pressure of water vapor must be subtracted from the overall pressure given. Table 5.2 reports pressure at 26°C (25.2 torr) and 28°C (28.3 torr), so take the average of the two values to obtain the partial pressure of water at 27°C. Volume must be in units of liters, pressure in atm, and temperature in kelvins. Once the moles of hydrogen produced are known, the molar ratio from the balanced equation is used to determine the moles of aluminum that reacted.

Solution:

$$V = 35.8 \text{ mL}$$

$$T = 27^\circ\text{C} + 273 = 300 \text{ K}$$

$$P_{\text{total}} = 751 \text{ mmHg}$$

$$n = \text{unknown}$$

$$P_{\text{water vapor}} = (28.3 + 25.2) \text{ torr}/2 = 26.75 \text{ torr} = 26.75 \text{ mmHg}$$

$$P_{\text{hydrogen}} = P_{\text{total}} - P_{\text{water vapor}} = 751 \text{ mmHg} - 26.75 \text{ mmHg} = 724.25 \text{ mmHg}$$

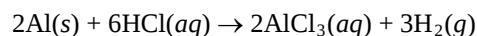
$$\text{Converting } P \text{ from mmHg to atm: } P = (724.25 \text{ mmHg}) \left( \frac{1 \text{ atm}}{760 \text{ mmHg}} \right) = 0.952960526 \text{ atm}$$

$$\text{Converting } V \text{ from mL to L: } V = (35.8 \text{ mL}) \left( \frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) = 0.0358 \text{ L}$$

$$PV = nRT$$

Solving for  $n$ :

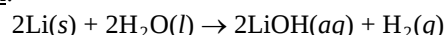
$$n = \frac{PV}{RT} = \frac{(0.952960526 \text{ atm})(0.0358 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(300. \text{ K})} = 0.0013851395 \text{ mol H}_2$$



$$\text{Mass (g) of Al} = (0.0013851395 \text{ mol H}_2) \left( \frac{2 \text{ mol Al}}{3 \text{ mol H}_2} \right) \left( \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} \right) = 0.024914 = \mathbf{0.0249 \text{ g Al}}$$

- 5.43 Plan: First, write the balanced equation. Convert mass of lithium to moles and use the molar ratio from the balanced equation to find the moles of hydrogen gas produced. Use the ideal gas law to find the volume of that amount of hydrogen. The problem specifies that the hydrogen gas is collected over water, so the partial pressure of water vapor must be subtracted from the overall pressure given. Table 5.2 reports the vapor pressure of water at 18°C (15.5 torr). Pressure must be in units of atm and temperature in kelvins.

Solution:



$$\text{Moles H}_2 = (0.84 \text{ g Li}) \left( \frac{1 \text{ mol Li}}{6.941 \text{ g Li}} \right) \left( \frac{1 \text{ mol H}_2}{2 \text{ mol Li}} \right) = 0.0605100 \text{ mol H}_2$$

Finding the volume of  $\text{H}_2$ :

$V$  = unknown

$$T = 18^\circ\text{C} + 273 = 291 \text{ K}$$

$$P_{\text{total}} = 725 \text{ mmHg}$$

$$n = 0.0605100 \text{ mol}$$

$$P_{\text{water vapor}} = 15.5 \text{ torr} = 15.5 \text{ mmHg}$$

$$P_{\text{hydrogen}} = P_{\text{total}} - P_{\text{water vapor}} = 725 \text{ mmHg} - 15.5 \text{ mmHg} = 709.5 \text{ mmHg}$$

$$\text{Converting } P \text{ from mmHg to atm: } P = (709.5 \text{ mmHg}) \left( \frac{1 \text{ atm}}{760 \text{ mmHg}} \right) = 0.933552631 \text{ atm}$$

$$PV = nRT$$

Solving for  $V$ :

$$V = \frac{nRT}{P} = \frac{(0.0605100 \text{ mol}) \left( 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (291 \text{ K})}{(0.933552631 \text{ atm})} = 1.5485 = \mathbf{1.5 \text{ L H}_2}$$

- 5.44 Plan: Rearrange the ideal gas law to calculate the density of the air from its molar mass. Temperature must be converted to kelvins and pressure to atmospheres.

Solution:

$$P = 744 \text{ torr}$$

$$T = 17^\circ\text{C} + 273 = 290 \text{ K}$$

or

$$T = 60^\circ\text{C} + 273 = 333 \text{ K}$$

$$\mathcal{M} \text{ of air} = 28.8 \text{ g/mol}$$

$$d = \text{unknown}$$

$$\text{Converting } P \text{ from torr to atm: } P = (744 \text{ torr}) \left( \frac{1 \text{ atm}}{760 \text{ torr}} \right) = 0.978947368 \text{ atm}$$

$$PV = nRT$$

Rearranging to solve for density:

At 17°C

$$d = \frac{P\mathcal{M}}{RT} = \frac{(0.978947368 \text{ atm})(28.8 \text{ g/mol})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(290 \text{ K})} = 1.18416 = \mathbf{1.18 \text{ g/L}}$$

At 60.0°C

$$d = \frac{P\mathcal{M}}{RT} = \frac{(0.978947368 \text{ atm})(28.8 \text{ g/mol})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(333 \text{ K})} = 1.03125 = \mathbf{1.03 \text{ g/L}}$$

- 5.45 Plan: The problem gives the mass, volume, temperature, and pressure of a gas; rearrange the formula  $PV = (m/M)RT$  to solve for the molar mass of the gas. Temperature must be in Kelvin and pressure in atm. The problem also states that the gas is a hydrocarbon, which by, definition, contains only carbon and hydrogen atoms. We are also told that each molecule of the gas contains five carbon atoms so we can use this information and the calculated molar mass to find out how many hydrogen atoms are present and the formula of the compound.

Solution:

$$V = 0.204 \text{ L}$$

$$T = 101^\circ\text{C} + 273 = 374 \text{ K}$$

$$P = 767 \text{ torr}$$

$$m = 0.482 \text{ g}$$

$M$  = unknown

$$\text{Converting } P \text{ from torr to atm: } P = (767 \text{ torr}) \left( \frac{1 \text{ atm}}{760 \text{ torr}} \right) = 1.009210526 \text{ atm}$$

$$PV = \left( \frac{m}{M} \right) RT$$

Solving for molar mass,  $M$ :

$$M = \frac{mRT}{PV} = \frac{(0.482 \text{ g}) \left( 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (374 \text{ K})}{(1.009210526 \text{ atm})(0.204 \text{ L})} = 71.8869 \text{ g/mol}$$

The mass of the five carbon atoms accounts for  $[5(12 \text{ g/mol})] = 60 \text{ g/mol}$ ; thus, the hydrogen atoms must make up the difference  $(72 - 60) = 12 \text{ g/mol}$ . A value of  $12 \text{ g/mol}$  corresponds to  $12 \text{ H}$  atoms. (Since fractional atoms are not possible, rounding is acceptable.) Therefore, the molecular formula is  $\text{C}_5\text{H}_{12}$ .

- 5.46 Plan: Solve the ideal gas law for moles of air. Temperature must be in units of kelvins. Use Avogadro's number to convert moles of air to molecules of air. The percent composition can be used to find the number of molecules (or atoms) of each gas in that total number of molecules.

Solution:

$$V = 1.00 \text{ L}$$

$$T = 25^\circ\text{C} + 273 = 298 \text{ K}$$

$$P = 1.00 \text{ atm}$$

$n$  = unknown

$$PV = nRT$$

Solving for  $n$ :

$$\text{Moles of air} = n = \frac{PV}{RT} = \frac{(1.00 \text{ atm})(1.00 \text{ L})}{\left( 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (298 \text{ K})} = 0.040873382 \text{ mol}$$

Converting moles of air to molecules of air:

$$\text{Molecules of air} = (0.040873382 \text{ mol}) \left( \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} \right) = 2.461395 \times 10^{22} \text{ molecules}$$

$$\begin{aligned} \text{Molecules of N}_2 &= (2.461395 \times 10^{22} \text{ air molecules}) \left( \frac{78.08\% \text{ N}_2 \text{ molecules}}{100\% \text{ air}} \right) \\ &= 1.921857 \times 10^{22} = \mathbf{1.92 \times 10^{22} \text{ molecules N}_2} \end{aligned}$$

$$\begin{aligned} \text{Molecules of O}_2 &= (2.461395 \times 10^{22} \text{ air molecules}) \left( \frac{20.94\% \text{ O}_2 \text{ molecules}}{100\% \text{ air}} \right) \\ &= 5.154161 \times 10^{21} = \mathbf{5.15 \times 10^{21} \text{ molecules O}_2} \end{aligned}$$

$$\begin{aligned} \text{Molecules of CO}_2 &= (2.461395 \times 10^{22} \text{ air molecules}) \left( \frac{0.05\% \text{ CO}_2 \text{ molecules}}{100\% \text{ air}} \right) \\ &= 1.2306975 \times 10^{19} = \mathbf{1 \times 10^{19} \text{ molecules CO}_2} \end{aligned}$$

$$\begin{aligned} \text{Molecules of Ar} &= (2.461395 \times 10^{22} \text{ air molecules}) \left( \frac{0.93\% \text{ Ar molecules}}{100\% \text{ air}} \right) \\ &= 2.289097 \times 10^{20} = \mathbf{2.3 \times 10^{20} \text{ molecules Ar}} \end{aligned}$$

- 5.47 Plan: Since you have the pressure, volume, and temperature, use the ideal gas law to solve for  $n$ , the total moles of gas. Pressure must be in units of atmospheres and temperature in units of kelvins. The partial pressure of  $\text{SO}_2$  can be found by multiplying the total pressure by the volume fraction of  $\text{SO}_2$ .

Solution:

$$\begin{array}{ll} \text{a) } V = 21 \text{ L} & T = 45^\circ\text{C} + 273 = 318 \text{ K} \\ P = 850 \text{ torr} & n = \text{unknown} \end{array}$$

Converting  $P$  from torr to atm:  $P = (850 \text{ torr}) \left( \frac{1 \text{ atm}}{760 \text{ torr}} \right) = 1.118421053 \text{ atm}$

$$PV = nRT$$

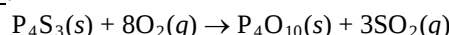
$$\text{Moles of gas} = n = \frac{PV}{RT} = \frac{(1.118421053 \text{ atm})(21 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(318 \text{ K})} = 0.89961 = \mathbf{0.90 \text{ mol gas}}$$

b) The equation  $P_{\text{SO}_2} = X_{\text{SO}_2} \times P_{\text{total}}$  can be used to find partial pressure. The information given in ppm is a way of expressing the proportion, or fraction, of  $\text{SO}_2$  present in the mixture. Since  $n$  is directly proportional to  $V$ , the volume fraction can be used in place of the mole fraction,  $X_{\text{SO}_2}$ . There are  $7.95 \times 10^3$  parts  $\text{SO}_2$  in a million parts of mixture, so volume fraction  $= (7.95 \times 10^3 / 1 \times 10^6) = 7.95 \times 10^{-3}$ .

$$P_{\text{D}_2} = \text{volume fraction} \times P_{\text{total}} = (7.95 \times 10^{-3})(850. \text{ torr}) = 6.7575 = \mathbf{6.76 \text{ torr}}$$

- 5.48 Plan: First, write the balanced equation. Convert mass of  $\text{P}_4\text{S}_3$  to moles and use the molar ratio from the balanced equation to find the moles of  $\text{SO}_2$  gas produced. Use the ideal gas law to find the volume of that amount of  $\text{SO}_2$ . Pressure must be in units of atm and temperature in kelvins.

Solution:



$$\text{Moles SO}_2 = (0.800 \text{ g P}_4\text{S}_3) \left( \frac{1 \text{ mol P}_4\text{S}_3}{220.09 \text{ g P}_4\text{S}_3} \right) \left( \frac{3 \text{ mol SO}_2}{1 \text{ mol P}_4\text{S}_3} \right) = 0.010905 \text{ mol SO}_2$$

Finding the volume of  $\text{SO}_2$ :

$$\begin{array}{ll} V = \text{unknown} & T = 32^\circ\text{C} + 273 = 305 \text{ K} \\ P = 725 \text{ torr} & n = 0.010905 \text{ mol} \end{array}$$

Converting  $P$  from torr to atm:  $P = (725 \text{ torr}) \left( \frac{1 \text{ atm}}{760 \text{ torr}} \right) = 0.953947368 \text{ atm}$

$$PV = nRT$$

Solving for  $V$ :

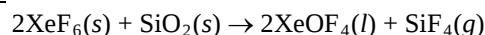
$$V = \frac{nRT}{P} = \frac{(0.010905 \text{ mol}) \left( 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (305 \text{ K})}{(0.953947368 \text{ atm})} = 0.28624918 \text{ L}$$

Converting  $V$  from L to mL:

$$V = (0.28624918 \text{ L}) \left( \frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) = 286.249 = \mathbf{286 \text{ mL SO}_2}$$

- 5.49 Plan: First, write the balanced equation. Given the amount of xenon hexafluoride that reacts, we can find the number of moles of silicon tetrafluoride gas formed by using the molar ratio in the balanced equation. Then, using the ideal gas law with the moles of gas, the temperature, and the volume, we can calculate the pressure of the silicon tetrafluoride gas. Temperature must be in units of kelvins.

Solution:



$$\text{Moles SiF}_4 = n = (2.00 \text{ g XeF}_6) \left( \frac{1 \text{ mol XeF}_6}{245.3 \text{ g XeF}_6} \right) \left( \frac{1 \text{ mol SiF}_4}{2 \text{ mol XeF}_6} \right) = 0.0040766 \text{ mol SiF}_4$$

Finding the pressure of  $\text{SiF}_4$ :

$V = 1.00 \text{ L}$   
 $P = \text{unknown}$   
 $PV = nRT$   
 Solving for  $P$ :

$T = 25^\circ\text{C} + 273 = 298 \text{ K}$   
 $n = 0.0040766 \text{ mol}$

$$\text{Pressure SiF}_4 = P = \frac{nRT}{V} = \frac{(0.0040766 \text{ mol SiF}_4) \left( 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (298 \text{ K})}{1.00 \text{ L}} = 0.099737 = \mathbf{0.0997 \text{ atm SiF}_4}$$

- 5.50 Plan: Use the ideal gas law with  $T$  and  $P$  constant; then volume is directly proportional to moles.

Solution:

$PV = nRT$ . At constant  $T$  and  $P$ ,  $V \propto n$ . Since the volume of the products has been decreased to  $\frac{1}{2}$  the original volume, the moles (and molecules) must have been decreased by a factor of  $\frac{1}{2}$  as well. **Cylinder A** best represents the products as there are 2 product molecules (there were 4 reactant molecules).

- 5.51 Plan: Write the balanced equation. Since the amounts of 2 reactants are given, this is a limiting reactant problem. To find the volume of  $\text{SO}_2$ , use the molar ratios from the balanced equation to find the number of moles of  $\text{SO}_2$  produced by each reactant. The smaller number of moles of product indicates the limiting reagent. Solve for moles of  $\text{SO}_2$  using the ideal gas law.

Solution:

Moles of oxygen:

$V = 228 \text{ L}$

$T = 220^\circ\text{C} + 273 = 493 \text{ K}$

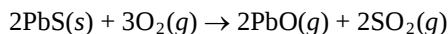
$P = 2 \text{ atm}$

$n = \text{unknown}$

$PV = nRT$

Solving for  $n$ :

$$\text{Moles of O}_2 = n = \frac{PV}{RT} = \frac{(2 \text{ atm})(228 \text{ L})}{\left( 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (493 \text{ K})} = 11.266 \text{ mol O}_2$$



$$\text{Moles SO}_2 \text{ from O}_2 = (11.266 \text{ mol O}_2) \left( \frac{2 \text{ mol SO}_2}{3 \text{ mol O}_2} \right) = 7.5107 \text{ mol SO}_2$$

$$\text{Moles SO}_2 \text{ from PbS} = (3.75 \text{ kg PbS}) \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left( \frac{1 \text{ mol PbS}}{239.3 \text{ g PbS}} \right) \left( \frac{2 \text{ mol SO}_2}{2 \text{ mol PbS}} \right) = 15.6707 \text{ mol SO}_2 \text{ (unrounded)}$$

$\text{O}_2$  is the limiting reagent because it forms less  $\text{SO}_2$ .

Finding the volume of  $\text{SO}_2$ :

$V = \text{unknown}$

$T = 0^\circ\text{C} + 273 = 273 \text{ K}$

$P_{\text{total}} = 1 \text{ atm}$

$n = 7.5107 \text{ mol}$

$PV = nRT$

Solving for  $V$ :

$$V = \frac{nRT}{P} = \frac{(7.5107 \text{ mol}) \left( 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (273 \text{ K})}{(1 \text{ atm})} = 168.34 = \mathbf{1.7 \times 10^2 \text{ L SO}_2}$$

- 5.52 As the temperature of the gas sample increases, the most probable speed increases. This will increase both the number of collisions per unit time and the force of each collision with the sample walls. Thus, the gas pressure increases.
- 5.53 At STP (or any identical temperature and pressure), the volume occupied by a mole of any gas will be identical. One mole of krypton has the same number of particles as one mole of helium and, at the same temperature, all of the gas particles have the same average kinetic energy, resulting in the same pressure and volume.

- 5.54 Plan: The molar masses of the three gases are 2.016 for H<sub>2</sub> (Flask A), 4.003 for He (Flask B), and 16.04 for CH<sub>4</sub> (Flask C). Since hydrogen has the smallest molar mass of the three gases, 4 g of H<sub>2</sub> will contain more gas molecules (about 2 mole's worth) than 4 g of He or 4 g of CH<sub>4</sub>. Since helium has a smaller molar mass than methane, 4 g of He will contain more gas molecules (about 1 mole's worth) than 4 g of CH<sub>4</sub> (about 0.25 mole's worth).
- Solution:
- a)  $P_A > P_B > P_C$  The pressure of a gas is proportional to the number of gas molecules ( $PV = nRT$ ). So, the gas sample with more gas molecules will have a greater pressure.
- b)  $E_A = E_B = E_C$  Average kinetic energy depends only on temperature. The temperature of each gas sample is 273 K, so they all have the same average kinetic energy.
- c)  $rate_A > rate_B > rate_C$  When comparing the speed of two gas molecules, the one with the lower mass travels faster.
- d)  $total\ E_A > total\ E_B > total\ E_C$  Since the average kinetic energy for each gas is the same (part b) of this problem), the total kinetic energy would equal the average times the number of molecules. Since the hydrogen flask contains the most molecules, its total kinetic energy will be the greatest.
- e)  $d_A = d_B = d_C$  Under the conditions stated in this problem, each sample has the same volume, 5 L, and the same mass, 4 g. Thus, the density of each is 4 g/5 L = 0.8 g/L.
- f) **Collision frequency (A) > collision frequency (B) > collision frequency (C)** The number of collisions depends on both the speed and the distance between gas molecules. Since hydrogen is the lightest molecule it has the greatest speed and the 5 L flask of hydrogen also contains the most molecules, so collisions will occur more frequently between hydrogen molecules than between helium molecules. By the same reasoning, collisions will occur more frequently between helium molecules than between methane molecules.
- 5.55 Plan: To find the ratio of effusion rates, calculate the inverse of the ratio of the square roots of the molar masses (Graham's law).
- Solution:
- $$\frac{\text{Rate H}_2}{\text{Rate UF}_6} = \sqrt{\frac{\text{molar mass UF}_6}{\text{molar mass H}_2}} = \sqrt{\frac{352.0 \text{ g/mol}}{2.016 \text{ g/mol}}} = 13.2137 = \mathbf{13.21}$$
- 5.56 Plan: To find the ratio of effusion rates, calculate the inverse of the ratio of the square roots of the molar masses (Graham's law).
- Solution:
- $$\frac{\text{Rate O}_2}{\text{Rate Kr}} = \sqrt{\frac{\text{molar mass Kr}}{\text{molar mass O}_2}} = \sqrt{\frac{83.80 \text{ g/mol}}{32.00 \text{ g/mol}}} = 1.618255 = \mathbf{1.618}$$
- 5.57 Plan: Recall that the heavier the gas, the slower the molecular speed. The molar mass of Ar is 39.95 g/mol while the molar mass of He is 4.003 g/mol.
- Solution:
- a) The gases have the same average kinetic energy because they are at the same temperature. The heavier Ar atoms are moving more slowly than the lighter He atoms to maintain the same average kinetic energy. Therefore, **Curve 1** with the lower average molecular speed, better represents the behavior of Ar.
- b) A gas that has a slower molecular speed would effuse more slowly, so **Curve 1** is the better choice.
- c) Fluorine gas exists as a diatomic molecule, F<sub>2</sub>, with  $\mathcal{M}$  = 38.00 g/mol. Therefore, F<sub>2</sub> is much closer in mass to Ar (39.95 g/mol) than He (4.003 g/mol), so **Curve 1** more closely represents the behavior of F<sub>2</sub>.
- 5.58 Plan: Recall that the lower the temperature, the lower the average kinetic energy and the slower the molecular speed.
- Solution:
- a) At the lower temperature, the average molecular speed is lower so **Curve 1** represents the gas at the lower temperature.
- b) When a gas has a higher kinetic energy, the molecules have a higher molecular speed. **Curve 2** with the larger average molecular speed represents the gas when it has a higher kinetic energy.
- c) If a gas has a higher diffusion rate, then the gas molecules are moving with a higher molecular speed as in **Curve 2**.

- 5.59 Plan: To find the ratio of effusion rates, calculate the inverse of the ratio of the square roots of the molar masses (Graham's law). Then use the ratio of effusion rates to find the time for the F<sub>2</sub> effusion. Effusion rate and time required for the effusion are inversely proportional.

Solution:

$$\begin{aligned} \mathcal{M} \text{ of He} &= 4.003 \text{ g/mol} & \mathcal{M} \text{ of F}_2 &= 38.00 \text{ g/mol} \\ \frac{\text{Rate He}}{\text{Rate F}_2} &= \sqrt{\frac{\text{molar mass F}_2}{\text{molar mass He}}} = \sqrt{\frac{38.00 \text{ g/mol}}{4.003 \text{ g/mol}}} = 3.08105 \text{ (unrounded)} \\ \frac{\text{Rate He}}{\text{Rate F}_2} &= \frac{\text{time F}_2}{\text{time He}} & \frac{3.08105}{1.00} &= \frac{\text{time F}_2}{4.85 \text{ min He}} & \text{Time F}_2 &= 14.9431 = \mathbf{14.9 \text{ min}} \end{aligned}$$

- 5.60 Plan: Effusion rate and time required for the effusion are inversely proportional. Therefore, time of effusion for a gas is directly proportional to the square root of its molar mass. The ratio of effusion times and the molar mass of H<sub>2</sub> are used to find the molar mass of the unknown gas.

Solution:

$$\mathcal{M} \text{ of H}_2 = 2.016 \text{ g/mol} \quad \text{Time of effusion of H}_2 = 2.42 \text{ min} \quad \text{Time of effusion of unknown} = 11.1 \text{ min}$$

$$\frac{\text{rate H}_2}{\text{rate unknown}} = \frac{\text{time unknown}}{\text{time H}_2} = \sqrt{\frac{\text{molar mass unknown}}{\text{molar mass H}_2}}$$

$$\frac{11.1 \text{ min}}{2.42 \text{ min}} = \sqrt{\frac{\text{molar mass unknown}}{2.016 \text{ g/mol}}}$$

$$4.586777 = \sqrt{\frac{\text{molar mass unknown}}{2.016 \text{ g/mol}}}$$

$$21.03852196 = \frac{\text{molar mass unknown}}{2.016 \text{ g/mol}}$$

$$\text{Molar mass unknown} = 42.41366 = \mathbf{42.4 \text{ g/mol}}$$

- 5.61 Plan: White phosphorus is a molecular form of the element phosphorus consisting of some number, x, of phosphorus atoms; the number of atoms in a molecule determines the molar mass of the phosphorus molecule. Use the relative rates of effusion of white phosphorus and neon (Graham's law) to determine the molar mass of white phosphorus. From the molar mass of white phosphorus, determine the number of phosphorus atoms, x, in one molecule of white phosphorus.

Solution:

$$\mathcal{M} \text{ of Ne} = 20.18 \text{ g/mol}$$

$$\frac{\text{Rate P}_x}{\text{Rate Ne}} = 0.404 = \sqrt{\frac{\text{molar mass Ne}}{\text{molar mass P}_x}}$$

$$0.404 = \sqrt{\frac{20.18 \text{ g/mol}}{\text{molar mass P}_x}}$$

$$(0.404)^2 = \frac{20.18 \text{ g/mol}}{\text{molar mass P}_x}$$

$$0.163216 = \frac{20.18 \text{ g/mol}}{\text{molar mass P}_x}$$

$$\text{Molar mass P}_x = 123.6398 \text{ g/mol}$$

$$\left( \frac{123.6398 \text{ g}}{\text{mol P}_x} \right) \left( \frac{1 \text{ mol P}}{30.97 \text{ g P}} \right) = 3.992244 = 4 \text{ mol P/mol P}_x \text{ or } 4 \text{ atoms P/molecule P}_x$$

Thus, **4 atoms per molecule**, so P<sub>x</sub> = P<sub>4</sub>.

- 5.62 Plan: Use the equation for root mean speed ( $u_{\text{rms}}$ ) to find this value for He at 0.°C and 30.°C and for Xe at 30.°C. The calculated root mean speed is then used in the kinetic energy equation to find the average kinetic energy for the two gases at 30.°C. Molar mass values must be in units of kg/mol and temperature in kelvins.  
Solution:

$$\text{a) } 0^{\circ}\text{C} = 273 \text{ K} \quad 30^{\circ}\text{C} + 273 = 303 \text{ K} \quad \mathcal{M} \text{ of He} = \left( \frac{4.003 \text{ g He}}{\text{mol}} \right) \left( \frac{1 \text{ kg}}{10^3 \text{ g}} \right) = 0.004003 \text{ kg/mol}$$

$$R = 8.314 \text{ J/mol}\cdot\text{K} \quad 1 \text{ J} = \text{kg}\cdot\text{m}^2/\text{s}^2$$

$$u_{\text{rms}} = \sqrt{\frac{3RT}{\mathcal{M}}}$$

$$u_{\text{rms}} \text{ He (at } 0^{\circ}\text{C)} = \sqrt{\frac{3 \left( 8.314 \frac{\text{J}}{\text{mol}\cdot\text{K}} \right) (273 \text{ K}) \left( \frac{\text{kg}\cdot\text{m}^2/\text{s}^2}{\text{J}} \right)}{0.004003 \text{ kg/mol}}} = 1.3042 \times 10^3 = \mathbf{1.30 \times 10^3 \text{ m/s}}$$

$$u_{\text{rms}} \text{ He (at } 30^{\circ}\text{C)} = \sqrt{\frac{3 \left( 8.314 \frac{\text{J}}{\text{mol}\cdot\text{K}} \right) (303 \text{ K}) \left( \frac{\text{kg}\cdot\text{m}^2/\text{s}^2}{\text{J}} \right)}{0.004003 \text{ kg/mol}}} = 1.3740 \times 10^3 = \mathbf{1.37 \times 10^3 \text{ m/s}}$$

$$\text{b) } 30^{\circ}\text{C} + 273 = 303 \text{ K} \quad \mathcal{M} \text{ of Xe} = \left( \frac{131.3 \text{ g Xe}}{\text{mol}} \right) \left( \frac{1 \text{ kg}}{10^3 \text{ g}} \right) = 0.1313 \text{ kg/mol}$$

$$R = 8.314 \text{ J/mol}\cdot\text{K} \quad 1 \text{ J} = \text{kg}\cdot\text{m}^2/\text{s}^2$$

$$u_{\text{rms}} = \sqrt{\frac{3RT}{\mathcal{M}}}$$

$$u_{\text{rms}} \text{ Xe (at } 30^{\circ}\text{C)} = \sqrt{\frac{3 \left( 8.314 \frac{\text{J}}{\text{mol}\cdot\text{K}} \right) (303 \text{ K}) \left( \frac{\text{kg}\cdot\text{m}^2/\text{s}^2}{\text{J}} \right)}{0.1313 \text{ kg/mol}}} = 239.913 \text{ m/s (unrounded)}$$

$$\text{Rate He/Rate Xe} = (1.3740 \times 10^3 \text{ m/s}) / (239.913 \text{ m/s}) = 5.727076 = \mathbf{5.73}$$

He molecules travel at almost 6 times the speed of Xe molecules.

$$\text{c) } E_k = \frac{1}{2} m u^2$$

$$E_{\text{He}} = \frac{1}{2} (0.004003 \text{ kg/mol}) (1.3740 \times 10^3 \text{ m/s})^2 \left( 1 \text{ J/kg}\cdot\text{m}^2/\text{s}^2 \right) = 3778.58 = \mathbf{3.78 \times 10^3 \text{ J/mol}}$$

$$E_{\text{Xe}} = \frac{1}{2} (0.1313 \text{ kg/mol}) (239.913 \text{ m/s})^2 \left( 1 \text{ J/kg}\cdot\text{m}^2/\text{s}^2 \right) = 3778.70 = \mathbf{3.78 \times 10^3 \text{ J/mol}}$$

$$\text{d) } \left( \frac{3778.58 \text{ J}}{\text{mol}} \right) \left( \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ atoms}} \right) = 6.2746 \times 10^{-21} = \mathbf{6.27 \times 10^{-21} \text{ J/He atom}}$$

- 5.63 Plan: Use Graham's law: the rate of effusion of a gas is inversely proportional to the square root of the molar mass. When comparing the speed of gas molecules, the one with the lowest mass travels the fastest.

Solution:

$$\text{a) } \mathcal{M} \text{ of } \text{S}_2\text{F}_2 = 102.14 \text{ g/mol}; \mathcal{M} \text{ of } \text{N}_2\text{F}_4 = 104.02 \text{ g/mol}; \mathcal{M} \text{ of } \text{SF}_4 = 108.07 \text{ g/mol}$$

$\text{SF}_4$  has the largest molar mass and  $\text{S}_2\text{F}_2$  has the smallest molar mass:  $\text{rate}_{\text{SF}_4} < \text{rate}_{\text{N}_2\text{F}_4} < \text{rate}_{\text{S}_2\text{F}_2}$

$$\text{b) } \frac{\text{Rate}_{\text{S}_2\text{F}_2}}{\text{Rate}_{\text{N}_2\text{F}_4}} = \sqrt{\frac{\text{molar mass } \text{N}_2\text{F}_4}{\text{molar mass } \text{S}_2\text{F}_2}} = \sqrt{\frac{104.02 \text{ g/mol}}{102.14 \text{ g/mol}}} = 1.009161 = \mathbf{1.0092:1}$$

$$\text{c) } \frac{\text{Rate X}}{\text{Rate } \text{SF}_4} = 0.935 = \sqrt{\frac{\text{molar mass } \text{SF}_4}{\text{molar mass X}}}$$

$$0.935 = \sqrt{\frac{108.07 \text{ g/mol}}{\text{molar mass X}}}$$

$$(0.935)^2 = \frac{108.07 \text{ g/mol}}{\text{molar mass X}}$$

$$0.874225 = \frac{108.07 \text{ g/mol}}{\text{molar mass X}}$$

$$\text{Molar mass X} = 123.61806 = \mathbf{124 \text{ g/mol}}$$

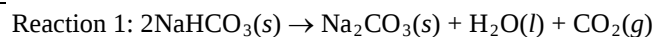
- 5.64 Interparticle attractions cause the real pressure to be *less than* ideal pressure, so it causes a **negative** deviation. The size of the interparticle attraction is related to the constant  $a$ . According to Table 5.4,  $a_{\text{N}_2} = 1.39$ ,  $a_{\text{Kr}} = 2.32$ , and  $a_{\text{CO}_2} = 3.59$ . Therefore,  $\text{CO}_2$  experiences a greater negative deviation in pressure than the other two gases:  $\text{N}_2 < \text{Kr} < \text{CO}_2$ .
- 5.65 Particle volume causes a positive deviation from ideal behavior. Thus,  $V_{\text{Real Gases}} > V_{\text{Ideal Gases}}$ . The particle volume is related to the constant  $b$ . According to Table 5.4,  $b_{\text{H}_2} = 0.0266$ ,  $b_{\text{O}_2} = 0.0318$ , and  $b_{\text{Cl}_2} = 0.0562$ . Therefore, the order is  $\text{H}_2 < \text{O}_2 < \text{Cl}_2$ .
- 5.66 Nitrogen gas behaves more ideally at **1 atm** than at 500 atm because at lower pressures the gas molecules are farther apart. An ideal gas is defined as consisting of gas molecules that act independently of the other gas molecules. When gas molecules are far apart they act more ideally, because intermolecular attractions are less important and the volume of the molecules is a smaller fraction of the container volume.
- 5.67  $\text{SF}_6$  behaves more ideally at **150°C**. At higher temperatures, intermolecular attractions become less important and the volume occupied by the molecules becomes less important.
- 5.68 Plan: Use the ideal gas law to find the number of moles of  $\text{O}_2$ . Moles of  $\text{O}_2$  is divided by 4 to find moles of Hb since  $\text{O}_2$  combines with Hb in a 4:1 ratio. Divide the given mass of Hb by the number of moles of Hb to obtain molar mass, g/mol. Temperature must be in units of kelvins, pressure in atm, and volume in L.  
Solution:  
 $V = 1.53 \text{ mL}$   
 $P = 743 \text{ torr}$   
 $T = 37^\circ\text{C} + 273 = 310 \text{ K}$   
 $n = \text{unknown}$   
 Converting  $V$  from mL to L:  $V = (1.53 \text{ mL}) \left( \frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) = 1.53 \times 10^{-3} \text{ L}$   
 Converting  $P$  from torr to atm:  $P = (743 \text{ torr}) \left( \frac{1 \text{ atm}}{760 \text{ torr}} \right) = 0.977631578 \text{ atm}$   
 $PV = nRT$   
 Solving for  $n$ :  

$$\text{Moles of O}_2 = n = \frac{PV}{RT} = \frac{(0.977631578 \text{ atm})(1.53 \times 10^{-3} \text{ L})}{\left( 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (310 \text{ K})} = 5.87708 \times 10^{-5} \text{ mol O}_2$$
  

$$\text{Moles Hb} = (5.87708 \times 10^{-5} \text{ mol O}_2) \left( \frac{1 \text{ mol Hb}}{4 \text{ mol O}_2} \right) = 1.46927 \times 10^{-5} \text{ mol Hb (unrounded)}$$
  

$$\text{Molar mass hemoglobin} = \frac{1.00 \text{ g Hb}}{1.46927 \times 10^{-5} \text{ Hb}} = 6.806098 \times 10^4 = \mathbf{6.81 \times 10^4 \text{ g/mol}}$$
- 5.69 Plan: First, write the balanced equations. Convert mass of  $\text{NaHCO}_3$  to moles and use the molar ratio from each balanced equation to find the moles of  $\text{CO}_2$  gas produced. Use the ideal gas law to find the volume of that amount of  $\text{CO}_2$ . Temperature must be in kelvins.

Solution:



$$\text{Moles CO}_2 = (1.00 \text{ g NaHCO}_3) \left( \frac{1 \text{ mol NaHCO}_3}{84.01 \text{ g NaHCO}_3} \right) \left( \frac{1 \text{ mol CO}_2}{2 \text{ mol NaHCO}_3} \right) = 5.95167 \times 10^{-3} \text{ mol CO}_2$$

Finding the volume of  $\text{CO}_2$ :

$V$  = unknown

$P = 0.975 \text{ atm}$

$PV = nRT$

Solving for  $V$ :

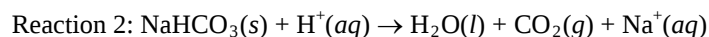
$$T = 200.^\circ\text{C} + 273 = 473 \text{ K}$$

$$n = 5.95167 \times 10^{-3} \text{ mol}$$

$$\text{Volume of CO}_2 = V = \frac{nRT}{P} = \frac{(5.95167 \times 10^{-3} \text{ mol}) \left( 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (473 \text{ K})}{(0.975 \text{ atm})} = 0.237049 \text{ L}$$

Converting  $V$  from L to mL:

$$V = (0.237049 \text{ L}) \left( \frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) = 237.049 = \mathbf{237 \text{ mL CO}_2 \text{ in Reaction 1}}$$



$$\text{Moles CO}_2 = (1.00 \text{ g NaHCO}_3) \left( \frac{1 \text{ mol NaHCO}_3}{84.01 \text{ g NaHCO}_3} \right) \left( \frac{1 \text{ mol CO}_2}{1 \text{ mol NaHCO}_3} \right) = 1.1903 \times 10^{-2} \text{ mol CO}_2$$

Finding the volume of  $\text{CO}_2$ :

$V$  = unknown

$P = 0.975 \text{ atm}$

$PV = nRT$

Solving for  $V$ :

$$T = 200.^\circ\text{C} + 273 = 473 \text{ K}$$

$$n = 1.1903 \times 10^{-2} \text{ mol}$$

$$\text{Volume of CO}_2 = V = \frac{nRT}{P} = \frac{(1.1903 \times 10^{-2} \text{ mol}) \left( 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (473 \text{ K})}{(0.975 \text{ atm})} = 0.4740986 \text{ L}$$

Converting  $V$  from L to mL:

$$V = (0.4740986 \text{ L}) \left( \frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) = 474.0986 = \mathbf{474 \text{ mL CO}_2 \text{ in Reaction 2}}$$

- 5.70 Plan: Convert the mass of  $\text{Cl}_2$  to moles and use the ideal gas law and van der Waals equation to find the pressure of the gas.

Solution:

$$\text{a) Moles Cl}_2: (0.5950 \text{ kg Cl}_2) \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left( \frac{1 \text{ mol Cl}_2}{70.90 \text{ g Cl}_2} \right) = 8.3921016 \text{ mol}$$

$V = 15.50 \text{ L}$

$n = 8.3921016 \text{ mol}$

Ideal gas law:  $PV = nRT$

Solving for  $P$ :

$$T = 225^\circ\text{C} + 273 = 498 \text{ K}$$

$P$  = unknown

$$P_{\text{IGL}} = \frac{nRT}{V} = \frac{(8.3921016 \text{ mol}) \left( 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (498 \text{ K})}{15.50 \text{ L}} = 22.1366 = \mathbf{22.1 \text{ atm}}$$

$$\text{b) van der Waals equation: } \left( P + \frac{n^2 a}{V^2} \right) (V - nb) = nRT$$

Solving for  $P$ :

$$P_{VDW} = \frac{nRT}{V - nb} - \frac{n^2 a}{V^2} \quad \text{From Table 5.4: } a = 6.49 \frac{\text{atm} \cdot \text{L}^2}{\text{mol}^2}; \quad b = 0.0562 \frac{\text{L}}{\text{mol}}$$

$n = 8.3921016 \text{ mol}$  from part a)

$$P_{VDW} = \frac{(8.3921016 \text{ mol Cl}_2) \left( 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (498 \text{ K}) - (8.3921016 \text{ mol Cl}_2)^2 \left( 6.49 \frac{\text{atm} \cdot \text{L}^2}{\text{mol}^2} \right)}{15.50 \text{ L} - (8.3921016 \text{ mol Cl}_2) \left( 0.0562 \frac{\text{L}}{\text{mol}} \right) - (15.50 \text{ L})^2}$$

$$= 20.928855 = \mathbf{20.9 \text{ atm}}$$

- 5.71 Plan: Rearrange the formula  $PV = (m/M)RT$  to solve for molar mass. Convert the volume in mL to L. Temperature must be in Kelvin. To find the molecular formulas of I, II, III, and IV, assume 100 g of each sample so the percentages are numerically equivalent to the masses of each element. Convert each of the masses to moles by using the molar mass of each element involved. Divide all moles by the lowest number of moles and convert to whole numbers to determine the empirical formula. The empirical formula mass and the calculated molar mass will then relate the empirical formula to the molecular formula. For gas IV, use Graham's law to find the molar mass

Solution:

a)  $V = 750.0 \text{ mL}$

$m = 0.1000 \text{ g}$

$M$  = unknown

$T = 70.00^\circ\text{C} + 273.15 = 343.15 \text{ K}$

$P = 0.05951 \text{ atm (I); } 0.07045 \text{ atm (II); } 0.05767 \text{ atm (III)}$

Converting  $V$  from mL to L:  $V = (750.0 \text{ mL}) \left( \frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) = 0.7500 \text{ L}$

$$PV = \left( \frac{m}{M} \right) RT$$

Solving for molar mass,  $M$ :

$$\text{Molar mass I} = M = \frac{mRT}{PV} = \frac{(0.1000 \text{ g}) \left( 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (343.15 \text{ K})}{(0.05951 \text{ atm})(0.7500 \text{ L})} = 63.0905 = \mathbf{63.09 \text{ g I/mol}}$$

$$\text{Molar mass II} = M = \frac{mRT}{PV} = \frac{(0.1000 \text{ g}) \left( 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (343.15 \text{ K})}{(0.07045 \text{ atm})(0.7500 \text{ L})} = 53.293 = \mathbf{53.29 \text{ g II/mol}}$$

$$\text{Molar mass III} = M = \frac{mRT}{PV} = \frac{(0.1000 \text{ g}) \left( 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (343.15 \text{ K})}{(0.05767 \text{ atm})(0.7500 \text{ L})} = 65.10349 = \mathbf{65.10 \text{ g III/mol}}$$

b) % H in I =  $100\% - 85.63\% = 14.37\% \text{ H}$

% H in II =  $100\% - 81.10\% = 18.90\% \text{ H}$

% H in III =  $100\% - 82.98\% = 17.02\% \text{ H}$

Assume 100 g of each so the mass percentages are also the grams of the element.

I

$$\text{Moles B} = (85.63 \text{ g B}) \left( \frac{1 \text{ mol B}}{10.81 \text{ g B}} \right) = 7.921369 \text{ mol B (unrounded)}$$

$$\text{Moles H} = (14.37 \text{ g H}) \left( \frac{1 \text{ mol H}}{1.008 \text{ g H}} \right) = 14.25595 \text{ mol H (unrounded)}$$

$$\left( \frac{7.921369 \text{ mol B}}{7.921369 \text{ mol B}} \right) = 1.00 \quad \left( \frac{14.25595 \text{ mol H}}{7.921369 \text{ mol B}} \right) = 1.7997$$

The hydrogen value is not close enough to a whole number to round. Thus, both amounts need to be multiplied by the smallest value to get near whole numbers. This value is 5. Multiplying each value by 5

gives  $(1.00 \times 5) = 5$  for B and  $(1.7997 \times 5) = 9$  for H. The empirical formula is  $B_5H_9$ , which has a formula mass of 63.12 g/mol. The empirical formula mass is near the molecular mass from part a) (63.09 g/mol). Therefore, the empirical and molecular formulas are both  **$B_5H_9$** .

II

$$\text{Moles B} = (81.10 \text{ g B}) \left( \frac{1 \text{ mol B}}{10.81 \text{ g B}} \right) = 7.50231 \text{ mol B (unrounded)}$$

$$\text{Moles H} = (18.90 \text{ g H}) \left( \frac{1 \text{ mol H}}{1.008 \text{ g H}} \right) = 18.750 \text{ mol H (unrounded)}$$

$$\left( \frac{7.50231 \text{ mol B}}{7.50231 \text{ mol B}} \right) = 1.00 \quad \left( \frac{18.750 \text{ mol H}}{7.50231 \text{ mol B}} \right) = 2.4992$$

The hydrogen value is not close enough to a whole number to round. Thus, both amounts need to be multiplied by the smallest value to get near whole numbers. This value is 2. Multiplying each value by 2 gives  $(1.00 \times 2) = 2$  for B and  $(2.4992 \times 2) = 5$  for H. The empirical formula is  $B_2H_5$ , which has a formula mass of 26.66 g/mol. Dividing the molecular formula mass from part a) by the empirical formula mass gives the relationship between the formulas:  $(53.29 \text{ g/mol}) / (26.66 \text{ g/mol}) = 2$ . The molecular formula is two times the empirical formula, or  **$B_4H_{10}$** .

III

$$\text{Moles B} = (82.98 \text{ g B}) \left( \frac{1 \text{ mol B}}{10.81 \text{ g B}} \right) = 7.6762 \text{ mol B (unrounded)}$$

$$\text{Moles H} = (17.02 \text{ g H}) \left( \frac{1 \text{ mol H}}{1.008 \text{ g H}} \right) = 16.8849 \text{ mol H (unrounded)}$$

$$\left( \frac{7.6762 \text{ mol B}}{7.6762 \text{ mol B}} \right) = 1.00 \quad \left( \frac{16.8849 \text{ mol H}}{7.6762 \text{ mol B}} \right) = 2.2$$

The hydrogen value is not close enough to a whole number to round. Thus, both amounts need to be multiplied by the smallest value to get near whole numbers. This value is 5. Multiplying each value by 5 gives  $(1.00 \times 5) = 5$  for B and  $(2.2 \times 5) = 11$  for H. The empirical formula is  $B_5H_{11}$ , which has a formula mass of 65.14 g/mol. The empirical formula mass is near the molecular mass from part a). Therefore, the empirical and molecular formulas are both  **$B_5H_{11}$** .

c)

$$\frac{\text{Rate SO}_2}{\text{Rate IV}} = \sqrt{\frac{\text{molar mass IV}}{\text{molar mass SO}_2}}$$

$$\left( \frac{250.0 \text{ mL}}{13.04 \text{ min}} \right) \left( \frac{350.0 \text{ mL}}{12.00 \text{ min}} \right) = 0.657318 = \sqrt{\frac{\text{molar mass IV}}{64.07 \text{ g/mol}}}$$

$$0.657318^2 = \frac{\text{molar mass IV}}{64.07 \text{ g/mol}}$$

$$\text{Molar mass IV} = 27.6825 = \mathbf{27.68 \text{ g/mol}}$$

$$\% \text{ H in IV} = 100\% - 78.14\% = 21.86\% \text{ H}$$

$$\text{Moles B} = (78.14 \text{ g B}) \left( \frac{1 \text{ mol B}}{10.81 \text{ g B}} \right) = 7.22849 \text{ mol B (unrounded)}$$

$$\text{Moles H} = (21.86 \text{ g H}) \left( \frac{1 \text{ mol H}}{1.008 \text{ g H}} \right) = 21.6865 \text{ mol H (unrounded)}$$

$$\left( \frac{7.22849 \text{ mol B}}{7.22849 \text{ mol B}} \right) = 1.00 \quad \left( \frac{21.6865 \text{ mol H}}{7.22849 \text{ mol B}} \right) = 3.00$$

The empirical formula is  $\text{BH}_3$ , which has a formula mass of 13.83 g/mol. Dividing the molecular formula mass by the empirical formula mass gives the relationship between the formulas:  $(27.68 \text{ g/mol})/(13.83 \text{ g/mol}) = 2$ . The molecular formula is two times the empirical formula, or  $\text{B}_2\text{H}_6$ .

- 5.72 Plan: Calculate the mole fraction of each gas; the partial pressure of each gas is directly proportional to its mole fraction so the gas with the highest mole fraction has the highest partial pressure. Remember that kinetic energy is directly proportional to Kelvin temperature.

Solution:

$$\text{a) } X_A = \frac{n_A}{n_{\text{total}}}$$

$$\text{I. } X_A = \frac{3 \text{ A particles}}{9 \text{ total particles}} = 0.33; \quad \text{II. } X_A = \frac{4 \text{ A particles}}{12 \text{ total particles}} = 0.33; \quad \text{III. } X_A = \frac{5 \text{ A particles}}{15 \text{ total particles}} = 0.33$$

The partial pressure of A is **the same** in all 3 samples since the mole fraction of A is the same in all samples.

$$\text{b) I. } X_B = \frac{3 \text{ B particles}}{9 \text{ total particles}} = 0.33; \quad \text{II. } X_B = \frac{3 \text{ B particles}}{12 \text{ total particles}} = 0.25; \quad \text{III. } X_B = \frac{3 \text{ B particles}}{15 \text{ total particles}} = 0.20$$

The partial pressure of B is lowest in **Sample III** since the mole fraction of B is the smallest in that sample.

c) All samples are at the same temperature,  $T$ , so all have the **same** average kinetic energy.

- 5.73 Plan: Partial pressures and mole fractions are calculated from Dalton's law of partial pressures:  $P_A = X_A \times P_{\text{total}}$ . Remember that 1 atm = 760 torr. Solve the ideal gas law for moles and then convert to molecules using Avogadro's number to calculate the number of  $\text{O}_2$  molecules in the volume of an average breath.

Solution:

a) Convert each mole percent to a mole fraction by dividing by 100%.  $P_{\text{total}} = 1 \text{ atm} = 760 \text{ torr}$

$$P_{\text{Nitrogen}} = X_{\text{Nitrogen}} \times P_{\text{total}} = 0.786 \times 760 \text{ torr} = 597.36 = \mathbf{597 \text{ torr N}_2}$$

$$P_{\text{Oxygen}} = X_{\text{Oxygen}} \times P_{\text{total}} = 0.209 \times 760 \text{ torr} = 158.84 = \mathbf{159 \text{ torr O}_2}$$

$$P_{\text{Carbon Dioxide}} = X_{\text{Carbon Dioxide}} \times P_{\text{total}} = 0.0004 \times 760 \text{ torr} = 0.304 = \mathbf{0.3 \text{ torr CO}_2}$$

$$P_{\text{Water}} = X_{\text{Water}} \times P_{\text{total}} = 0.0046 \times 760 \text{ torr} = 3.496 = \mathbf{3.5 \text{ torr H}_2\text{O}}$$

b) Mole fractions can be calculated by rearranging Dalton's law of partial pressures:

$$X_A = \frac{P_A}{P_{\text{total}}} \text{ and multiply by 100 to express mole fraction as percent}$$

$$P_{\text{total}} = 1 \text{ atm} = 760 \text{ torr}$$

$$X_{\text{Nitrogen}} = \frac{569 \text{ torr}}{760 \text{ torr}} \times 100\% = 74.8684 = \mathbf{74.9 \text{ mol\% N}_2}$$

$$X_{\text{Oxygen}} = \frac{104 \text{ torr}}{760 \text{ torr}} \times 100\% = 13.6842 = \mathbf{13.7 \text{ mol\% O}_2}$$

$$X_{\text{Carbon Dioxide}} = \frac{40 \text{ torr}}{760 \text{ torr}} \times 100\% = 5.263 = \mathbf{5.3 \text{ mol\% CO}_2}$$

$$X_{\text{Water}} = \frac{47 \text{ torr}}{760 \text{ torr}} \times 100\% = 6.1842 = \mathbf{6.2 \text{ mol\% H}_2\text{O}}$$

$$\text{c) } V = 0.50 \text{ L}$$

$$P = 104 \text{ torr}$$

$$T = 37^\circ\text{C} + 273 = 310 \text{ K}$$

$$n = \text{unknown}$$

$$\text{Converting } P \text{ from torr to atm: } P = (104 \text{ torr}) \left( \frac{1 \text{ atm}}{760 \text{ torr}} \right) = 0.136842105 \text{ atm}$$

$$PV = nRT$$

Solving for  $n$ :

$$n = \frac{PV}{RT} = \frac{(0.136842105 \text{ atm})(0.50 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(310 \text{ K})} = 0.0026883 \text{ mol O}_2$$

$$\begin{aligned}\text{Molecules of O}_2 &= (0.0026883 \text{ mol O}_2) \left( \frac{6.022 \times 10^{23} \text{ molecules O}_2}{1 \text{ mol O}_2} \right) \\ &= 1.6189 \times 10^{21} = \mathbf{1.6 \times 10^{21} \text{ molecules O}_2}\end{aligned}$$

- 5.74 Plan: Convert the mass of Ra to moles and then atoms using Avogadro's number. Convert from number of Ra atoms to Rn atoms produced per second and then to Rn atoms produced per day. The number of Rn atoms is converted to moles and then the ideal gas law is used to find the volume of this amount of Rn.

Solution:

$$\text{Atoms Ra} = (1.0 \text{ g Ra}) \left( \frac{1 \text{ mol Ra}}{226 \text{ g Ra}} \right) \left( \frac{6.022 \times 10^{23} \text{ Ra atoms}}{1 \text{ mol Ra}} \right) = 2.664602 \times 10^{21} \text{ Ra atoms}$$

$$\text{Atoms Rn produced/s} = (2.664602 \times 10^{21} \text{ Ra atoms}) \left( \frac{1.373 \times 10^4 \text{ Rn atoms}}{1.0 \times 10^{15} \text{ Ra atoms}} \right) = 3.65849855 \times 10^{10} \text{ Rn atoms/s}$$

$$\begin{aligned}\text{Moles Rn produced/day} &= \left( \frac{3.65849855 \times 10^{10} \text{ Rn atoms}}{\text{s}} \right) \left( \frac{3600 \text{ s}}{\text{h}} \right) \left( \frac{24 \text{ h}}{\text{day}} \right) \left( \frac{1 \text{ mol Rn}}{6.022 \times 10^{23} \text{ Rn atoms}} \right) \\ &= 5.248992 \times 10^{-9} \text{ mole Rn/day}\end{aligned}$$

$$PV = nRT$$

Solving for V (at STP):

$$\begin{aligned}\text{Volume of Rn} = V &= \frac{nRT}{P} = \frac{(5.248992 \times 10^{-9} \text{ mol}) \left( 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (273 \text{ K})}{(1 \text{ atm})} \\ &= 1.17647 \times 10^{-7} = \mathbf{1.2 \times 10^{-7} \text{ L Rn}}\end{aligned}$$

- 5.75 Plan: For part a), since the volume, temperature, and pressure of the gas are changing, use the combined gas law. For part b), use the ideal gas law to solve for moles of air and then moles of N<sub>2</sub>.

Solution:

$$\begin{array}{ll}\text{a) } P_1 = 1450. \text{ mmHg} & P_2 = 1 \text{ atm} \\ V_1 = 208 \text{ mL} & V_2 = \text{unknown} \\ T_1 = 286 \text{ K} & T_2 = 298 \text{ K}\end{array}$$

$$\text{Converting } P_1 \text{ from mmHg to atm: } P_1 = (1450. \text{ mmHg}) \left( \frac{1 \text{ atm}}{760 \text{ mmHg}} \right) = 1.9079 \text{ atm}$$

Arranging the ideal gas law and solving for V<sub>2</sub>:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$V_2 = V_1 \left( \frac{T_2}{T_1} \right) \left( \frac{P_1}{P_2} \right) = (208 \text{ L}) \left( \frac{298 \text{ K}}{286 \text{ K}} \right) \left( \frac{1.9079 \text{ atm}}{1 \text{ atm}} \right) = 413.494 \text{ mL} = \mathbf{4 \times 10^2 \text{ mL}}$$

$$\begin{array}{ll}\text{b) } V = 208 \text{ mL} & T = 286 \text{ K} \\ P = 1450 \text{ mmHg} = 1.9079 \text{ atm} & n = \text{unknown}\end{array}$$

$$\text{Converting } V \text{ from mL to L: } V = (208 \text{ mL}) \left( \frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) = 0.208 \text{ L}$$

$$PV = nRT$$

Solving for n:

$$\text{Moles of air} = n = \frac{PV}{RT} = \frac{(1.9079 \text{ atm})(0.208 \text{ L})}{\left( 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (286 \text{ K})} = 0.016901 \text{ mol air}$$

$$\text{Mole of N}_2 = (0.016901 \text{ mol}) \left( \frac{77\% \text{ N}_2}{100\%} \right) = 0.01301 = \mathbf{0.013 \text{ mol N}_2}$$

- 5.76 Plan: The amounts of both reactants are given, so the first step is to identify the limiting reactant. Write the balanced equation and use molar ratios to find the number of moles of NO<sub>2</sub> produced by each reactant. The smaller number of moles of product indicates the limiting reagent. Solve for volume of NO<sub>2</sub> using the ideal gas law.

Solution:



$$\text{Moles NO}_2 \text{ from Cu} = (4.95 \text{ cm}^3) \left( \frac{8.95 \text{ g Cu}}{\text{cm}^3} \right) \left( \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} \right) \left( \frac{2 \text{ mol NO}_2}{1 \text{ mol Cu}} \right) = 1.394256 \text{ mol NO}_2$$

$$\begin{aligned} \text{Moles NO}_2 \text{ from HNO}_3 &= (230.0 \text{ mL}) \left( \frac{68.0\% \text{ HNO}_3}{100\%} \right) \left( \frac{1 \text{ cm}^3}{1 \text{ mL}} \right) \left( \frac{1.42 \text{ g}}{\text{cm}^3} \right) \left( \frac{1 \text{ mol HNO}_3}{63.02 \text{ g}} \right) \left( \frac{2 \text{ mol NO}_2}{4 \text{ mol HNO}_3} \right) \\ &= 1.7620 \text{ mol NO}_2 \end{aligned}$$

Since less product can be made from the copper, it is the limiting reactant and excess nitric acid will be left after the reaction goes to completion. Use the calculated number of moles of NO<sub>2</sub> and the given temperature and pressure in the ideal gas law to find the volume of nitrogen dioxide produced. Note that nitrogen dioxide is the only gas involved in the reaction.

$V = \text{unknown}$

$$T = 28.2^\circ\text{C} + 273.2 = 301.4 \text{ K}$$

$P = 735 \text{ torr}$

$$n = 1.394256 \text{ mol NO}_2$$

$$\text{Converting } P \text{ from torr to atm: } P = (735 \text{ torr}) \left( \frac{1 \text{ atm}}{760 \text{ torr}} \right) = 0.967105 \text{ atm}$$

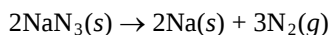
$$PV = nRT$$

Solving for  $V$ :

$$V = \frac{nRT}{P} = \frac{(1.394256 \text{ mol}) \left( 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (301.4 \text{ K})}{(0.967105 \text{ atm})} = 35.67429 = \mathbf{35.7 \text{ L NO}_2}$$

- 5.77 Plan: First, write the balanced equation. Convert mass of NaN<sub>3</sub> to moles and use the molar ratio from the balanced equation to find the moles of nitrogen gas produced. Use the ideal gas law to find the volume of that amount of nitrogen. The problem specifies that the nitrogen gas is collected over water, so the partial pressure of water vapor must be subtracted from the overall pressure given. Table 5.2 reports the vapor pressure of water at 26°C (25.2 torr). Pressure must be in units of atm and temperature in kelvins.

Solution:



$$\text{Moles N}_2 = (50.0 \text{ g NaN}_3) \left( \frac{1 \text{ mol NaN}_3}{65.02 \text{ g NaN}_3} \right) \left( \frac{3 \text{ mol N}_2}{2 \text{ mol NaN}_3} \right) = 1.15349 \text{ mol N}_2$$

Finding the volume of N<sub>2</sub>:

$V = \text{unknown}$

$$T = 26^\circ\text{C} + 273 = 299 \text{ K}$$

$P_{\text{total}} = 745.5 \text{ mmHg}$

$$n = 1.15319 \text{ mol}$$

$P_{\text{water vapor}} = 25.2 \text{ torr} = 25.2 \text{ mmHg}$

$$P_{\text{nitrogen}} = P_{\text{total}} - P_{\text{water vapor}} = 745.5 \text{ mmHg} - 25.2 \text{ mmHg} = 720.3 \text{ mmHg}$$

$$\text{Converting } P \text{ from mmHg to atm: } P = (720.3 \text{ mmHg}) \left( \frac{1 \text{ atm}}{760 \text{ mmHg}} \right) = 0.9477632 \text{ atm}$$

$$PV = nRT$$

Solving for  $V$ :

$$V = \frac{nRT}{P} = \frac{(1.15349 \text{ mol}) \left( 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (299 \text{ K})}{(0.9477632 \text{ atm})} = 29.8764 = \mathbf{29.9 \text{ L N}_2}$$

- 5.78 Plan: Use the percent composition information to find the empirical formula of the compound. Assume 100 g of sample so the percentages are numerically equivalent to the masses of each element. Convert each of the masses to moles by using the molar mass of each element involved. Divide all moles by the lowest number of moles and convert to whole numbers to determine the empirical formula. Rearrange the formula  $PV = (m/\mathcal{M})RT$  to solve for molar mass. The empirical formula mass and the calculated molar mass will then relate the empirical formula to the molecular formula.

Solution:

Empirical formula:

Assume 100 g of each so the mass percentages are also the grams of the element.

$$\text{Moles C} = (64.81 \text{ g C}) \left( \frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) = 5.39634 \text{ mol C}$$

$$\text{Moles H} = (13.60 \text{ g H}) \left( \frac{1 \text{ mol H}}{1.008 \text{ g H}} \right) = 13.49206 \text{ mol H}$$

$$\text{Moles O} = (21.59 \text{ g O}) \left( \frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 1.349375 \text{ mol O}$$

$$\left( \frac{5.39634 \text{ mol C}}{1.349375 \text{ mol O}} \right) = 4 \quad \left( \frac{13.49206 \text{ mol H}}{1.349375 \text{ mol O}} \right) = 10 \quad \left( \frac{1.349375 \text{ mol O}}{1.349375 \text{ mol O}} \right) = 1.00$$

Empirical formula =  $\text{C}_4\text{H}_{10}\text{O}$  (empirical formula mass = 74.12 g/mol)

Molecular formula:

$$V = 2.00 \text{ mL}$$

$$T = 25^\circ\text{C} + 273 = 298 \text{ K}$$

$$m = 2.57 \text{ g}$$

$$P = 0.420 \text{ atm}$$

$\mathcal{M}$  = unknown

$$PV = \left( \frac{m}{\mathcal{M}} \right) RT$$

Solving for molar mass,  $\mathcal{M}$ :

$$\text{Molar mass} = \mathcal{M} = \frac{mRT}{PV} = \frac{(2.57 \text{ g}) \left( 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (298 \text{ K})}{(0.420 \text{ atm})(2.00 \text{ L})} = 74.85 \text{ g/mol}$$

Since the molar mass (74.85 g/mol) and the empirical formula mass (74.12 g/mol) are similar, the empirical and molecular formulas must both be:  **$\text{C}_4\text{H}_{10}\text{O}$**

- 5.79 Plan: The empirical formula for aluminum chloride is  $\text{AlCl}_3$  ( $\text{Al}^{3+}$  and  $\text{Cl}^-$ ). The empirical formula mass is (133.33 g/mol). Calculate the molar mass of the gaseous species from the ratio of effusion rates (Graham's law). This molar mass, divided by the empirical weight, should give a whole-number multiple that will yield the molecular formula.

Solution:

$$\frac{\text{Rate unknown}}{\text{Rate He}} = 0.122 = \sqrt{\frac{\text{molar mass He}}{\text{molar mass unknown}}}$$

$$0.122 = \sqrt{\frac{4.003 \text{ g/mol}}{\text{molar mass unknown}}}$$

$$0.014884 = \frac{4.003 \text{ g/mol}}{\text{molar mass unknown}}$$

$$\text{Molar mass unknown} = 268.9465 \text{ g/mol}$$

The whole-number multiple is  $268.9465/133.33$ , which is about 2. Therefore, the molecular formula of the gaseous species is  $2 \times (\text{AlCl}_3) = \text{Al}_2\text{Cl}_6$ .

- 5.80 Plan: First, write the balanced equation for the reaction:  $2\text{SO}_2 + \text{O}_2 \rightarrow 2\text{SO}_3$ . The total number of moles of gas will change as the reaction occurs since 3 moles of reactant gas forms 2 moles of product gas. From the volume, temperature, and pressures given, we can calculate the number of moles of gas before and after the reaction using the ideal gas law. For each mole of  $\text{SO}_3$  formed, the total number of moles of gas decreases by 1/2 mole. Thus, twice the decrease in moles of gas equals the moles of  $\text{SO}_3$  formed.

Solution:

Moles of gas before and after reaction:

$$V = 2.00 \text{ L}$$

$$T = 800. \text{ K}$$

$$P_{\text{total}} = 1.90 \text{ atm}$$

$$n = \text{unknown}$$

$$PV = nRT$$

$$\text{Initial moles} = n = \frac{PV}{RT} = \frac{(1.90 \text{ atm})(2.00 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(800. \text{ K})} = 0.05785627 \text{ mol}$$

$$\text{Final moles} = n = \frac{PV}{RT} = \frac{(1.65 \text{ atm})(2.00 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(800. \text{ K})} = 0.050243605 \text{ mol}$$

$$\begin{aligned} \text{Moles of } \text{SO}_3 \text{ produced} &= 2 \times \text{decrease in the total number of moles} \\ &= 2 \times (0.05785627 \text{ mol} - 0.050243605 \text{ mol}) \\ &= 0.01522533 = \mathbf{1.52 \times 10^{-2} \text{ mol}} \end{aligned}$$

Check: If the starting amount is 0.0578 total moles of  $\text{SO}_2$  and  $\text{O}_2$ , then  $x + y = 0.0578 \text{ mol}$ , where  $x = \text{mol of } \text{SO}_2$  and  $y = \text{mol of } \text{O}_2$ . After the reaction:

$$(x - z) + (y - 0.5z) + z = 0.0502 \text{ mol}$$

Where  $z = \text{mol of } \text{SO}_3 \text{ formed} = \text{mol of } \text{SO}_2 \text{ reacted} = 2(\text{mol of } \text{O}_2 \text{ reacted})$ .

Subtracting the two equations gives:

$$x - (x - z) + y - (y - 0.5z) - z = 0.0578 - 0.0502$$

$$z = 0.0152 \text{ mol } \text{SO}_3$$

The approach of setting up two equations and solving them gives the same result as above.

- 5.81 Plan: Use the density of  $\text{C}_2\text{H}_4$  to find the volume of one mole of gas. Then use the van der Waals equation with 1.00 mol of gas to find the pressure of the gas (the mole ratio is 1:1, so the number of moles of gas remains the same).

Solution:

$$\text{a) } (1 \text{ mole } \text{C}_2\text{H}_4) \left( \frac{28.05 \text{ g } \text{C}_2\text{H}_4}{1 \text{ mole } \text{C}_2\text{H}_4} \right) \left( \frac{1 \text{ mL}}{0.215 \text{ g}} \right) \left( \frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) = 0.130465 \text{ L} = 0.130 \text{ L}$$

$$V = 0.130 \text{ L}$$

$$T = 10^\circ\text{C} + 273 = 283 \text{ K} + 950 \text{ K} = 1233 \text{ K}$$

$$P_{\text{total}} = \text{unknown}$$

$$n = 1.00 \text{ mol}$$

$$\text{From Table 5.4 for } \text{CH}_4: a = 2.25 \frac{\text{atm}\cdot\text{L}^2}{\text{mol}^2}; \quad b = 0.0428 \frac{\text{L}}{\text{mol}}$$

$$\left( P + \frac{n^2 a}{V^2} \right) (V - nb) = nRT$$

$$\text{Pressure of } \text{CH}_4 = P_{\text{VDW}} = \frac{nRT}{V - nb} - \frac{n^2 a}{V^2}$$

$$P_{\text{VDW}} = \frac{(1.00 \text{ mol}) \left( 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (1233 \text{ K})}{0.130 \text{ L} - 1.00 \text{ mol} (0.0428 \text{ L/mol})} - \frac{(1.00 \text{ mol})^2 \left( 2.25 \frac{\text{atm}\cdot\text{L}^2}{\text{mol}^2} \right)}{(0.130 \text{ L})^2} = 1027.7504 = \mathbf{1028 \text{ atm}}$$

$$\text{b) } \frac{PV}{RT} = \frac{(1028 \text{ atm})(0.130 \text{ L})}{\left( 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (1233 \text{ K})} = \mathbf{1.32}$$

This value is smaller than that shown in Figure 5.23 for CH<sub>4</sub>. The temperature in this situation is very high (1233 K). At high temperatures, the gas particles have high kinetic energy. Thus the gas particles have the energy to overcome the effects of intermolecular attraction and the gas behaves more ideally.

- 5.82 Plan: Write a balanced equation for the reaction. Use the mole ratio of the product gases and the total pressure of the mixture to find the partial pressure of each gas. Use the ideal gas law with the partial pressure of either product gas to find the moles of that gas produced; the mole ratio between product gas and reactant allows calculation of the mass of original reactant.

Solution:

The reaction is:  $2 \text{NCl}_3(l) \xrightarrow{\Delta} \text{N}_2(g) + 3 \text{Cl}_2(g)$

The decomposition of all the NCl<sub>3</sub> means that the final pressure must be due to the N<sub>2</sub> and the Cl<sub>2</sub>. The product gases are present at a 1:3 ratio, and the total moles are 4.

a) Partial pressure N<sub>2</sub> = P<sub>nitrogen</sub> = X<sub>nitrogen</sub> P<sub>total</sub> = (1 mol N<sub>2</sub>/4 mol total) (754 mmHg)  
= 188.5 = **188 mmHg N<sub>2</sub>**

Partial pressure Cl<sub>2</sub> = P<sub>nitrogen</sub> = X<sub>nitrogen</sub> P<sub>total</sub> = (3 mol Cl<sub>2</sub>/4 mol total) (754 mmHg) = 565.5 = **566 mmHg Cl<sub>2</sub>**

b) The mass of NCl<sub>3</sub> may be determined several ways. Using the partial pressure of Cl<sub>2</sub> gives:

$$\text{Moles Cl}_2 = n = \frac{PV}{RT} = \frac{(565.5 \text{ mmHg})(2.50 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)((273 + 95) \text{ K})} \left(\frac{1 \text{ atm}}{760 \text{ mmHg}}\right) = 0.0615698 \text{ mol Cl}_2$$

$$\text{Mass NCl}_3 = (0.0615698 \text{ mol Cl}_2) \left(\frac{2 \text{ mol NCl}_3}{3 \text{ mol Cl}_2}\right) \left(\frac{120.36 \text{ g NCl}_3}{1 \text{ mol NCl}_3}\right) = 4.94036 = \mathbf{4.94 \text{ g NCl}_3}$$

- 5.83 Plan: Use the percent composition information to find the empirical formula of the compound. Assume 100 g of sample so the percentages are numerically equivalent to the masses of each element. Convert each of the masses to moles by using the molar mass of each element involved. Divide all moles by the lowest number of moles and convert to whole numbers to determine the empirical formula. Rearrange the formula  $PV = (m/\mathcal{M})RT$  to solve for molar mass. The empirical formula mass and the calculated molar mass will then relate the empirical formula to the molecular formula.

Solution:

Empirical formula:

Assume 100 g of each so the mass percentages are also the grams of the element.

$$\text{Moles Si} = (33.01 \text{ g Si}) \left(\frac{1 \text{ mol Si}}{28.09 \text{ g Si}}\right) = 1.17515 \text{ mol Si}$$

$$\text{Moles F} = (66.99 \text{ g F}) \left(\frac{1 \text{ mol F}}{19.00 \text{ g F}}\right) = 3.525789 \text{ mol F}$$

$$\left(\frac{1.17515 \text{ mol Si}}{1.17515 \text{ mol Si}}\right) = 1 \quad \left(\frac{3.525789 \text{ mol F}}{1.17515 \text{ mol Si}}\right) = 3$$

Empirical formula = SiF<sub>3</sub> (empirical formula mass = 85.1 g/mol)

Molecular formula:

V = 0.250 L

T = 27°C + 273 = 300 K

m = 2.60 g

P = 1.50 atm

$\mathcal{M}$  = unknown

$$PV = \left(\frac{m}{\mathcal{M}}\right)RT$$

Solving for molar mass,  $\mathcal{M}$ :

$$\text{Molar mass} = \mathcal{M} = \frac{mRT}{PV} = \frac{(2.60 \text{ g}) \left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right) (300 \text{ K})}{(1.50 \text{ atm})(0.250 \text{ L})} = 170.768 \text{ g/mol}$$

The molar mass (170.768 g/mol) is twice the empirical formula mass (85.1 g/mol), so the molecular formula must be twice the empirical formula, or  $2 \times \text{SiF}_3 = \text{Si}_2\text{F}_6$ .

- 5.84 **Plan:** Four moles of gas ( $\text{NH}_3$ ,  $\text{CO}$ ,  $\text{N}_2$ , and  $\text{HCNO}$ ) are formed from the decomposition of 1 mole of azodicarbonamide. Two of those moles of gas,  $\text{NH}_3$  and  $\text{HCNO}$ , further react to form solid nonvolatile polymers. So the decomposition of 1 mole of azodicarbonamide leads to the overall formation of two moles of gas. Convert the given mass of azodicarbonamide to moles and multiply by 2 to find the number of moles of gas produced. Use the ideal gas law to find the volume of that amount of gas at STP.

**Solution:**

$$\text{Moles of gas formed} = (1.00 \text{ g } \text{C}_2\text{H}_4\text{N}_4\text{O}_2) \left( \frac{1 \text{ mol } \text{C}_2\text{H}_4\text{N}_4\text{O}_2}{116.09 \text{ g } \text{C}_2\text{H}_4\text{N}_4\text{O}_2} \right) \left( \frac{2 \text{ moles of gas}}{1 \text{ mol } \text{C}_2\text{H}_4\text{N}_4\text{O}_2} \right) = 0.017228 \text{ mol}$$

$$\text{Volume (L) of gas} = V = \frac{nRT}{P} = \frac{(0.017228 \text{ mol}) \left( 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (273 \text{ K})}{(1.00 \text{ atm})} = 0.386136 \text{ L}$$

Converting  $V$  from L to mL:

$$V = (0.386136 \text{ L}) \left( \frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) = 386.136 = \mathbf{386 \text{ mL gas}}$$

- 5.85 **Plan:** Write a balanced reaction based on the information given about the volumes of gases produced. Since the volume of a gas is proportional to the number of moles of the gas we can equate volume and moles.

**Solution:**

a) A preliminary equation for this reaction is  $4\text{C}_x\text{H}_y\text{N}_z + n\text{O}_2 \rightarrow 4\text{CO}_2 + 2\text{N}_2 + 10\text{H}_2\text{O}$ .

Since the organic compound does not contain oxygen, the only source of oxygen as a reactant is oxygen gas. To form 4 volumes of  $\text{CO}_2$  would require 4 volumes of  $\text{O}_2$  and to form 10 volumes of  $\text{H}_2\text{O}$  would require 5 volumes of  $\text{O}_2$ . Thus, 9 **volumes of  $\text{O}_2$**  were required.

b) From a volume ratio of  $4\text{CO}_2:2\text{N}_2:10\text{H}_2\text{O}$  we deduce a mole ratio of  $4\text{C}:4\text{N}:20\text{H}$  or  $1\text{C}:1\text{N}:5\text{H}$  for an empirical formula of  **$\text{CH}_5\text{N}$** .

- 5.86 a) There is a total of  $6 \times 10^6$  blue particles and  $6 \times 10^6$  black particles. When equilibrium is reached after opening the stopcocks, the particles will be evenly distributed among the three containers. Therefore, container B will have  **$2 \times 10^6$  blue particles and  $2 \times 10^6$  black particles**.  
 b) The particles are evenly distributed so container A has  **$2 \times 10^6$  blue particles and  $2 \times 10^6$  black particles**.  
 c) There are  $2 \times 10^6$  blue particles and  $2 \times 10^6$  black particles in C for a total of  $4 \times 10^6$  particles.

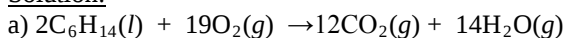
$$\text{Final pressure in C} = \left( 4 \times 10^6 \text{ particles} \right) \left( \frac{750 \text{ torr}}{6 \times 10^6 \text{ particles}} \right) = \mathbf{500 \text{ torr}}$$

d) There are  $2 \times 10^6$  blue particles and  $2 \times 10^6$  black particles in B for a total of  $4 \times 10^6$  particles.

$$\text{Final pressure in B} = \left( 4 \times 10^6 \text{ particles} \right) \left( \frac{750 \text{ torr}}{6 \times 10^6 \text{ particles}} \right) = \mathbf{500 \text{ torr}}$$

- 5.87 **Plan:** Write the balanced equation for the combustion of  $n$ -hexane. For part a), assuming a 1.00 L sample of air at STP, use the molar ratio in the balanced equation to find the volume of  $n$ -hexane required to react with the oxygen in 1.00 L of air. Convert the volume  $n$ -hexane to volume % and divide by 2 to obtain the LFL. For part b), use the LFL calculated in part a) to find the volume of  $n$ -hexane required to produce a flammable mixture and then use the ideal gas law to find moles of  $n$ -hexane. Convert moles of  $n$ -hexane to mass and then to volume using the density.

**Solution:**



For a 1.00 L sample of air at STP:

$$\text{Volume of } \text{C}_6\text{H}_{14} \text{ vapor needed} = (1.00 \text{ L air}) \left( \frac{20.9 \text{ L } \text{O}_2}{100 \text{ L air}} \right) \left( \frac{2 \text{ L } \text{C}_6\text{H}_{14}}{19 \text{ L } \text{O}_2} \right) = 0.0220 \text{ L } \text{C}_6\text{H}_{14}$$

$$\text{Volume \% of C}_6\text{H}_{14} = \frac{\text{C}_6\text{H}_{14} \text{ volume}}{\text{air volume}}(100) = \frac{0.0220 \text{ L C}_6\text{H}_{14}}{1.00 \text{ L air}}(100) = 2.2\% \text{ C}_6\text{H}_{14}$$

$$\text{LFL} = 0.5(2.2\%) = \mathbf{1.1\% \text{ C}_6\text{H}_{14}}$$

$$\text{b) Volume of C}_6\text{H}_{14} \text{ vapor} = (1.000 \text{ m}^3 \text{ air}) \left( \frac{1 \text{ L}}{10^{-3} \text{ m}^3} \right) \left( \frac{1.1\% \text{ C}_6\text{H}_{14}}{100\% \text{ air}} \right) = 11.0 \text{ L C}_6\text{H}_{14}$$

$$V = 11.0 \text{ L}$$

$$T = 0^\circ\text{C} + 273 = 273 \text{ K}$$

$$P = 1 \text{ atm}$$

$$n = \text{unknown}$$

$$PV = nRT$$

Solving for  $n$ :

$$\text{Moles of C}_6\text{H}_{14} = n = \frac{PV}{RT} = \frac{(1 \text{ atm})(11.0 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(273 \text{ K})} = 0.490780 \text{ mol C}_6\text{H}_{14}$$

$$\text{Volume of C}_6\text{H}_{14} \text{ liquid} = (0.490780 \text{ mol C}_6\text{H}_{14}) \left( \frac{86.17 \text{ g C}_6\text{H}_{14}}{1 \text{ mol C}_6\text{H}_{14}} \right) \left( \frac{1 \text{ mL}}{0.660 \text{ g C}_6\text{H}_{14}} \right) = 64.0765 = \mathbf{64 \text{ mL}}$$

### **C<sub>6</sub>H<sub>14</sub>**

- 5.88 Plan: To find the factor by which a diver's lungs would expand, find the factor by which  $P$  changes from 125 ft to the surface, and apply Boyle's law. To find that factor, calculate  $P_{\text{seawater}}$  at 125 ft by converting the given depth from ft-seawater to mmHg to atm and adding the surface pressure (1.00 atm).

Solution:

$$P(\text{H}_2\text{O}) = (125 \text{ ft}) \left( \frac{12 \text{ in}}{1 \text{ ft}} \right) \left( \frac{2.54 \text{ cm}}{1 \text{ in}} \right) \left( \frac{10^{-2} \text{ m}}{1 \text{ cm}} \right) \left( \frac{1 \text{ mm}}{10^{-3} \text{ m}} \right) = 3.81 \times 10^4 \text{ mmH}_2\text{O}$$

$$P(\text{Hg}): \frac{h_{\text{H}_2\text{O}}}{h_{\text{Hg}}} = \frac{d_{\text{Hg}}}{d_{\text{H}_2\text{O}}} \quad \frac{3.81 \times 10^4 \text{ mmH}_2\text{O}}{h_{\text{Hg}}} = \frac{13.5 \text{ g/mL}}{1.04 \text{ g/mL}} \quad h_{\text{Hg}} = 2935.1111 \text{ mmHg}$$

$$P(\text{Hg}) = (2935.1111 \text{ mmHg}) \left( \frac{1 \text{ atm}}{760 \text{ mm Hg}} \right) = 3.861988 \text{ atm}$$

$$P_{\text{total}} = (1.00 \text{ atm}) + (3.861988 \text{ atm}) = 4.861988 \text{ atm}$$

Use Boyle's law to find the volume change of the diver's lungs:

$$P_1 V_1 = P_2 V_2$$

$$\frac{V_2}{V_1} = \frac{P_1}{P_2} \quad \frac{V_2}{V_1} = \frac{4.861988 \text{ atm}}{1 \text{ atm}} = \mathbf{4.86}$$

To find the depth to which the diver could ascend safely, use the given safe expansion factor (1.5) and the pressure at 125 ft,  $P_{125}$ , to find the safest ascended pressure,  $P_{\text{safe}}$ .

$$P_{125}/P_{\text{safe}} = 1.5$$

$$P_{\text{safe}} = P_{125}/1.5 = (4.861988 \text{ atm})/1.5 = 3.241325 \text{ atm}$$

Convert the pressure in atm to pressure in ft of seawater using the conversion factors above. Subtract this distance from the initial depth to find how far the diver could ascend.

$$h(\text{Hg}): (4.861988 - 3.241325 \text{ atm}) \left( \frac{760 \text{ mmHg}}{1 \text{ atm}} \right) = 1231.7039 \text{ mmHg}$$

$$\frac{h_{\text{H}_2\text{O}}}{h_{\text{Hg}}} = \frac{d_{\text{Hg}}}{d_{\text{H}_2\text{O}}} \quad \frac{h_{\text{H}_2\text{O}}}{1231.7039 \text{ mmHg}} = \frac{13.5 \text{ g/mL}}{1.04 \text{ g/mL}} \quad h_{\text{H}_2\text{O}} = 15988.464 \text{ mm}$$

$$h_{\text{H}_2\text{O}} (\text{ft}) = (15988.464 \text{ mmH}_2\text{O}) \left( \frac{10^{-3} \text{ m}}{1 \text{ mm}} \right) \left( \frac{1.094 \text{ yd}}{1 \text{ m}} \right) \left( \frac{3 \text{ ft}}{1 \text{ yd}} \right) = 52.4741 \text{ ft}$$

Therefore, the diver can safely ascend 52.5 ft to a depth of  $(125 - 52.4741) = 72.5259 = \mathbf{73 \text{ ft}}$ .

- 5.89 Plan: The moles of gas may be found using the ideal gas law. Multiply moles of gas by Avogadro's number to obtain the number of molecules.

Solution:

$$V = 1 \text{ mL} = 0.001 \text{ L}$$

$$T = 500 \text{ K}$$

$$P = 10^{-8} \text{ mmHg}$$

$$n = \text{unknown}$$

$$\text{Converting } P \text{ from mmHg to atm: } P = (10^{-8} \text{ mmHg}) \left( \frac{1 \text{ atm}}{760 \text{ mmHg}} \right) = 1.315789 \times 10^{-11} \text{ atm}$$

$$PV = nRT$$

Solving for  $n$ :

$$\text{Moles of gas} = n = \frac{PV}{RT} = \frac{(1.315789 \times 10^{-11} \text{ atm})(0.001 \text{ L})}{\left(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(500 \text{ K})} = 3.2053337 \times 10^{-16} \text{ mol gas}$$

$$\text{Molecules} = (3.2053337 \times 10^{-16} \text{ mol}) \left( \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} \right) = 1.93025 \times 10^8 = \mathbf{10^8 \text{ molecules}}$$

(The  $10^{-8}$  mmHg limits the significant figures.)

- 5.90 Plan: Use the equation for root mean speed ( $u_{\text{rms}}$ ) to find this value for  $\text{O}_2$  at  $0^\circ\text{C}$ . Molar mass values must be in units of kg/mol and temperature in kelvins. Divide the root mean speed by the mean free path to obtain the collision frequency.

Solution:

$$\text{a) } 0^\circ\text{C} = 273 \text{ K} \quad \mathcal{M} \text{ of } \text{O}_2 = \left( \frac{32.00 \text{ g O}_2}{\text{mol}} \right) \left( \frac{1 \text{ kg}}{10^3 \text{ g}} \right) = 0.03200 \text{ kg/mol}$$

$$R = 8.314 \text{ J/mol} \cdot \text{K} \quad 1 \text{ J} = \text{kg} \cdot \text{m}^2/\text{s}^2$$

$$u_{\text{rms}} = \sqrt{\frac{3RT}{\mathcal{M}}}$$

$$u_{\text{rms}} = \sqrt{\frac{3 \left( 8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}} \right) (273 \text{ K})}{0.03200 \text{ kg/mol}} \left( \frac{\text{kg} \cdot \text{m}^2/\text{s}^2}{\text{J}} \right)} = 461.2878 = \mathbf{461 \text{ m/s}}$$

$$\text{b) Collision frequency} = \frac{u_{\text{rms}}}{\text{mean free path}} = \frac{461.2878 \text{ m/s}}{6.33 \times 10^{-8} \text{ m}} = 7.2873 \times 10^9 = \mathbf{7.29 \times 10^9 \text{ s}^{-1}}$$

- 5.91 Plan: Use the ideal gas law to calculate the molar volume, the volume of exactly one mole of gas, at the temperature and pressure given in the problem.

Solution:

$$V = \text{unknown}$$

$$T = 730. \text{ K}$$

$$P = 90 \text{ atm}$$

$$n = 1.00 \text{ mol}$$

$$PV = nRT$$

Solving for  $V$ :

$$V = \frac{nRT}{P} = \frac{(1.00 \text{ mol}) \left( 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (730. \text{ K})}{(90 \text{ atm})} = 0.66592 = \mathbf{0.67 \text{ L/mol}}$$

- 5.92 Plan: Use the ideal gas law to determine the total moles of gas produced. The total moles multiplied by the fraction of each gas gives the moles of that gas which may be converted to metric tons.

Solution:

$$V = 1.5 \times 10^3 \text{ m}^3$$

$$T = 298 \text{ K}$$

$$P = 1 \text{ atm}$$

$$n = \text{unknown}$$

$$\text{Converting } V \text{ from m}^3 \text{ to L: } V = (1.5 \times 10^3 \text{ m}^3) \left( \frac{1 \text{ L}}{10^{-3} \text{ m}^3} \right) = 1.5 \times 10^6 \text{ L}$$

$$PV = nRT$$

Solving for  $n$ :

$$\text{Moles of gas/day} = n = \frac{PV}{RT} = \frac{(1 \text{ atm})(1.5 \times 10^6 \text{ L})}{\left(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(298 \text{ K})} = 6.13101 \times 10^5 \text{ mol/day}$$

$$\text{Moles of gas/yr} = \left( \frac{6.13101 \times 10^5 \text{ mol}}{\text{day}} \right) \left( \frac{365.25 \text{ day}}{1 \text{ yr}} \right) = 2.23935 \times 10^7 \text{ mol/yr}$$

$$\text{Mass CO}_2 = (0.4896) \left( \frac{2.23935 \times 10^7 \text{ mol}}{\text{yr}} \right) \left( \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \right) \left( \frac{1 \text{ kg}}{10^3 \text{ g}} \right) \left( \frac{1 \text{ t}}{10^3 \text{ kg}} \right) = 482.519 = \mathbf{4.83 \times 10^2 \text{ t CO}_2/\text{yr}}$$

$$\text{Mass CO} = (0.0146) \left( \frac{2.23935 \times 10^7 \text{ mol}}{\text{yr}} \right) \left( \frac{28.01 \text{ g CO}}{1 \text{ mol CO}} \right) \left( \frac{1 \text{ kg}}{10^3 \text{ g}} \right) \left( \frac{1 \text{ t}}{10^3 \text{ kg}} \right) = 9.15773 = \mathbf{9.16 \text{ t CO/yr}}$$

$$\text{Mass H}_2\text{O} = (0.3710) \left( \frac{2.23935 \times 10^7 \text{ mol}}{\text{yr}} \right) \left( \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) \left( \frac{1 \text{ kg}}{10^3 \text{ g}} \right) \left( \frac{1 \text{ t}}{10^3 \text{ kg}} \right) = 149.70995 = \mathbf{1.50 \times 10^2 \text{ t H}_2\text{O/yr}}$$

$$\text{Mass SO}_2 = (0.1185) \left( \frac{2.23935 \times 10^7 \text{ mol}}{\text{yr}} \right) \left( \frac{64.07 \text{ g SO}_2}{1 \text{ mol SO}_2} \right) \left( \frac{1 \text{ kg}}{10^3 \text{ g}} \right) \left( \frac{1 \text{ t}}{10^3 \text{ kg}} \right) = 170.018 = \mathbf{1.70 \times 10^2 \text{ t SO}_2/\text{yr}}$$

$$\text{Mass S}_2 = (0.0003) \left( \frac{2.23935 \times 10^7 \text{ mol}}{\text{yr}} \right) \left( \frac{64.14 \text{ g S}_2}{1 \text{ mol S}_2} \right) \left( \frac{1 \text{ kg}}{10^3 \text{ g}} \right) \left( \frac{1 \text{ t}}{10^3 \text{ kg}} \right) = 0.4308957 = \mathbf{4 \times 10^{-1} \text{ t S}_2/\text{yr}}$$

$$\text{Mass H}_2 = (0.0047) \left( \frac{2.23935 \times 10^7 \text{ mol}}{\text{yr}} \right) \left( \frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} \right) \left( \frac{1 \text{ kg}}{10^3 \text{ g}} \right) \left( \frac{1 \text{ t}}{10^3 \text{ kg}} \right) = 0.21218 = \mathbf{2.1 \times 10^{-1} \text{ t H}_2/\text{yr}}$$

$$\text{Mass HCl} = (0.0008) \left( \frac{2.23935 \times 10^7 \text{ mol}}{\text{yr}} \right) \left( \frac{36.46 \text{ g HCl}}{1 \text{ mol HCl}} \right) \left( \frac{1 \text{ kg}}{10^3 \text{ g}} \right) \left( \frac{1 \text{ t}}{10^3 \text{ kg}} \right) = 0.6531736 = \mathbf{6 \times 10^{-1} \text{ t HCl/yr}}$$

$$\text{Mass H}_2\text{S} = (0.0003) \left( \frac{2.23935 \times 10^7 \text{ mol}}{\text{yr}} \right) \left( \frac{34.09 \text{ g H}_2\text{S}}{1 \text{ mol H}_2\text{S}} \right) \left( \frac{1 \text{ kg}}{10^3 \text{ g}} \right) \left( \frac{1 \text{ t}}{10^3 \text{ kg}} \right) = 0.229018 = \mathbf{2 \times 10^{-1} \text{ t H}_2\text{S/yr}}$$

- 5.93 Plan: Use the molar ratio from the balanced equation to find the moles of H<sub>2</sub> and O<sub>2</sub> required to form 28.0 moles of water. Then use the ideal gas law in part a) and van der Waals equation in part b) to find the pressure needed to provide that number of moles of each gas.

Solution:

a) The balanced chemical equation is:  $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$

$$\text{Moles H}_2 = (28.0 \text{ mol H}_2\text{O}) \left( \frac{2 \text{ mol H}_2}{2 \text{ mol H}_2\text{O}} \right) = 28.0 \text{ mol H}_2$$

$$\text{Moles O}_2 = (28.0 \text{ mol H}_2\text{O}) \left( \frac{1 \text{ mol O}_2}{2 \text{ mol H}_2\text{O}} \right) = 14.0 \text{ mol O}_2$$

$$V = 20.0 \text{ L}$$

$$T = 23.8^\circ\text{C} + 273.2 = 297 \text{ K}$$

$$P = \text{unknown}$$

$$n = 28.0 \text{ mol H}_2; 14.0 \text{ mol O}_2$$

$$PV = nRT$$

Solving for  $P$ :

$$P_{\text{IGL of H}_2} = \frac{nRT}{V} = \frac{(28.0 \text{ mol}) \left( 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (297 \text{ K})}{20.0 \text{ L}} = 34.137 = \mathbf{34.1 \text{ atm H}_2}$$

$$P_{\text{IGL of O}_2} = \frac{nRT}{V} = \frac{(14.0 \text{ mol}) \left( 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (297 \text{ K})}{20.0 \text{ L}} = 17.06859 = \mathbf{17.1 \text{ atm O}_2}$$

b)  $V = 20.0 \text{ L}$

$P = \text{unknown}$

Van der Waals constants from Table 5.4:

$$\text{H}_2: a = 0.244 \frac{\text{atm} \cdot \text{L}^2}{\text{mol}^2}; \quad b = 0.0266 \frac{\text{L}}{\text{mol}}$$

$$\text{O}_2: a = 1.36 \frac{\text{atm} \cdot \text{L}^2}{\text{mol}^2}; \quad b = 0.0318 \frac{\text{L}}{\text{mol}}$$

$$\left( P + \frac{n^2 a}{V^2} \right) (V - nb) = nRT$$

$$P_{\text{VDW}} = \frac{nRT}{V - nb} - \frac{n^2 a}{V^2}$$

$$P_{\text{VDW of H}_2} = \frac{(28.0 \text{ mol}) \left( 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (297 \text{ K})}{20.0 \text{ L} - 28.0 \text{ mol} (0.0266 \text{ L/mol})} - \frac{(28.0 \text{ mol})^2 \left( 0.244 \frac{\text{atm} \cdot \text{L}^2}{\text{mol}^2} \right)}{(20.0 \text{ L})^2} = 34.9631 = \mathbf{35.0 \text{ atm H}_2}$$

$$P_{\text{VDW of O}_2} = \frac{(14.0 \text{ mol}) \left( 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (297 \text{ K})}{20.0 \text{ L} - 28.0 \text{ mol} (0.0318 \text{ L/mol})} - \frac{(14.0 \text{ mol})^2 \left( 1.36 \frac{\text{atm} \cdot \text{L}^2}{\text{mol}^2} \right)}{(20.0 \text{ L})^2} = 16.78228 = \mathbf{16.8 \text{ atm O}_2}$$

c) The van der Waals value for hydrogen is slightly higher than the value from the ideal gas law. The van der Waals value for oxygen is slightly lower than the value from the ideal gas law.

- 5.94 Plan: Use the molarity and volume of the solution to find the moles of HBr needed to make the solution. Then use the ideal gas law to find the volume of that number of moles of HBr gas at the given conditions.

Solution:

$$\text{Moles of HBr in the hydrobromic acid: } \left( \frac{1.20 \text{ mol HBr}}{\text{L}} \right) (3.50 \text{ L}) = 4.20 \text{ mol HBr}$$

$V = \text{unknown}$

$P = 0.965 \text{ atm}$

$PV = nRT$

Solving for  $V$ :

$$V = \frac{nRT}{P} = \frac{(4.20 \text{ mol}) \left( 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (302 \text{ K})}{(0.965 \text{ atm})} = 107.9126 = \mathbf{108 \text{ L HBr}}$$

- 5.95 Plan:  $V$  and  $T$  are not given, so the ideal gas law cannot be used. The total pressure of the mixture is given. Use  $P_A = X_A \times P_{\text{total}}$  to find the mole fraction of each gas and then the mass fraction. The total mass of the two gases is 35.0 g.

Solution:

$$P_{\text{total}} = P_{\text{krypton}} + P_{\text{carbon dioxide}} = 0.708 \text{ atm}$$

The NaOH absorbed the  $\text{CO}_2$  leaving the Kr, thus  $P_{\text{krypton}} = 0.250 \text{ atm}$ .

$$P_{\text{carbon dioxide}} = P_{\text{total}} - P_{\text{krypton}} = 0.708 \text{ atm} - 0.250 \text{ atm} = 0.458 \text{ atm}$$

Determining mole fractions:  $P_A = X_A \times P_{\text{total}}$

$$\text{Carbon dioxide: } X = \frac{P_{\text{CO}_2}}{P_{\text{total}}} = \frac{0.458 \text{ atm}}{0.708 \text{ atm}} = 0.64689$$

$$\text{Krypton: } X = \frac{P_{\text{Kr}}}{P_{\text{total}}} = \frac{0.250 \text{ atm}}{0.708 \text{ atm}} = 0.353107$$

$$\text{Relative mass fraction} = \left[ \frac{(0.353107) \left( \frac{83.80 \text{ g Kr}}{\text{mol}} \right)}{(0.64689) \left( \frac{44.01 \text{ g CO}_2}{\text{mol}} \right)} \right] = 1.039366$$

$$35.0 \text{ g} = x \text{ g CO}_2 + (1.039366 x) \text{ g Kr}$$

$$35.0 \text{ g} = 2.039366 x$$

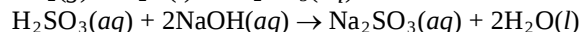
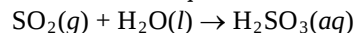
$$\text{Grams CO}_2 = x = (35.0 \text{ g}) / (2.039366) = 17.16219581 = \mathbf{17.2 \text{ g CO}_2}$$

$$\text{Grams Kr} = 35.0 \text{ g} - 17.162 \text{ g CO}_2 = 17.83780419 = \mathbf{17.8 \text{ g Kr}}$$

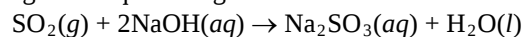
- 5.96 Plan: Write the balanced equations. Use the ideal gas law to find the moles of SO<sub>2</sub> gas and then use the molar ratio between SO<sub>2</sub> and NaOH to find moles and then molarity of the NaOH solution.

Solution:

The balanced chemical equations are:



Combining these equations gives:



$$V = 0.200 \text{ L}$$

$$T = 19^\circ\text{C} + 273 = 292 \text{ K}$$

$$P = 745 \text{ mmHg}$$

$$n = \text{unknown}$$

$$\text{Converting } P \text{ from mmHg to atm: } P = (745 \text{ mmHg}) \left( \frac{1 \text{ atm}}{760 \text{ mmHg}} \right) = 0.980263 \text{ atm}$$

$$PV = nRT$$

Solving for  $n$ :

$$\text{Moles of SO}_2 = n = \frac{PV}{RT} = \frac{(0.980263 \text{ atm})(0.200 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(292 \text{ K})} = 8.17799 \times 10^{-3} \text{ mol SO}_2$$

$$\text{Moles of NaOH} = (8.17799 \times 10^{-3} \text{ mol SO}_2) \left( \frac{2 \text{ mol NaOH}}{1 \text{ mol SO}_2} \right) = 0.01635598 \text{ mol NaOH}$$

$$M \text{ NaOH} = \frac{\text{mol NaOH}}{\text{volume of NaOH}} = \frac{0.01635598 \text{ mol NaOH}}{10.0 \text{ mL}} \left( \frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) = 1.635598 = \mathbf{1.64 \text{ M NaOH}}$$

- 5.97 Plan: Use the ideal gas law to find the number of moles of CO<sub>2</sub> and H<sub>2</sub>O in part a). The molar mass is then used to convert moles to mass. Temperature must be in units of kelvins, pressure in atm, and volume in L. For part b), use the molar ratio in the balanced equation to find the moles and then mass of C<sub>6</sub>H<sub>12</sub>O<sub>6</sub> that produces the number of moles of CO<sub>2</sub> exhaled during 8 h.

Solution:

$$\text{a) } V = 300 \text{ L}$$

$$T = 37.0^\circ\text{C} + 273.2 = 310.2 \text{ K}$$

$$P = 30.0 \text{ torr}$$

$$n = \text{unknown}$$

$$\text{Converting } P \text{ from torr to atm: } P = (30.0 \text{ torr}) \left( \frac{1 \text{ atm}}{760 \text{ torr}} \right) = 0.0394737 \text{ atm}$$

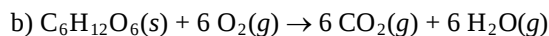
$$PV = nRT$$

Solving for  $n$ :

$$\text{Moles of CO}_2 = \text{moles of H}_2\text{O} = n = \frac{PV}{RT} = \frac{(0.0394737 \text{ atm})(300 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(310.2 \text{ K})} = 0.464991 \text{ mol}$$

$$\text{Mass (g) of CO}_2 = (0.464991 \text{ mol CO}_2) \left( \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \right) = 20.4643 = \mathbf{20.5 \text{ g CO}_2}$$

$$\text{Mass (g) of H}_2\text{O} = (0.464991 \text{ mol H}_2\text{O}) \left( \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) = 8.3791 = \mathbf{8.38 \text{ g H}_2\text{O}}$$



$$\text{Moles of CO}_2 \text{ exhaled in 8 h} = \left( \frac{0.464991 \text{ mol CO}_2}{\text{h}} \right) (8 \text{ h}) = 3.719928 \text{ mol CO}_2$$

$$\begin{aligned} \text{Mass (g) of C}_6\text{H}_{12}\text{O}_6 &= (3.719928 \text{ mol CO}_2) \left( \frac{1 \text{ mol C}_6\text{H}_{12}\text{O}_6}{6 \text{ mol CO}_2} \right) \left( \frac{180.16 \text{ g C}_6\text{H}_{12}\text{O}_6}{1 \text{ mol C}_6\text{H}_{12}\text{O}_6} \right) \\ &= 111.6970 = 1 \times 10^2 \text{ g C}_6\text{H}_{12}\text{O}_6 (= \text{body mass lost}) \end{aligned}$$

(This assumes the significant figures are limited by the 8 h.)

5.98 a) Derive  $u_{\text{rms}} = \sqrt{\frac{3RT}{\mathcal{M}}}$

Set the given relationships equal to each other.

$$\frac{1}{2} m \overline{u^2} = \frac{3}{2} \left( \frac{R}{N_A} \right) T \quad \text{Multiply each side by 2 and divide by } m.$$

$$\overline{u^2} = \frac{3}{m} \left( \frac{R}{N_A} \right) T$$

$$\overline{u^2} = \frac{3RT}{mN_A}$$

Solve for  $u$  by taking the square root of each side; substitute molar mass,  $\mathcal{M}$ , for  $mN_A$  (mass of one molecule x Avogadro's number of molecules).

$$u_{\text{rms}} = \sqrt{\frac{3RT}{\mathcal{M}}}$$

b) Derive Graham's Law  $\frac{\sqrt{\mathcal{M}_1}}{\sqrt{\mathcal{M}_2}} = \frac{\text{rate}_2}{\text{rate}_1}$

At a given  $T$ , the average kinetic energy is equal for two substances, with molecular masses  $m_1$  and  $m_2$ :

$$\overline{E_k} = \frac{1}{2} m_1 \overline{u_1^2} = \frac{1}{2} m_2 \overline{u_2^2}$$

$$m_1 \overline{u_1^2} = m_2 \overline{u_2^2}$$

$$\frac{m_1}{m_2} = \frac{\overline{u_2^2}}{\overline{u_1^2}} \rightarrow \frac{\sqrt{m_1}}{\sqrt{m_2}} = \frac{\overline{u_2}}{\overline{u_1}}$$

The average molecular speed,  $u$ , is directly proportional to the rate of effusion. Therefore, substitute "rate" for each " $u$ ." In addition, the molecular mass is directly proportional to the molar mass, so substitute  $\mathcal{M}$  for each  $m$ :

$$\frac{\sqrt{\mathcal{M}_1}}{\sqrt{\mathcal{M}_2}} = \frac{\text{rate}_2}{\text{rate}_1}$$

- 5.99 Plan: Use the ideal gas law to find the moles of gas occupying the tank at 85% of the 85.0 atm ranking. Then use van der Waals equation to find the pressure of this number of moles of gas.

Solution:

a)  $V = 850. \text{ L}$

$T = 298 \text{ K}$

$$P = (85.0 \text{ atm}) \left( \frac{80\%}{100\%} \right) = 68.0 \text{ atm} \quad n = \text{unknown}$$

$$PV = nRT$$

Solving for  $n$ :

$$\text{Moles of Cl}_2 = n = \frac{PV}{RT} = \frac{(68.0 \text{ atm})(850. \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(298 \text{ K})} = 2.36248 \times 10^3 = \mathbf{2.36 \times 10^3 \text{ mol Cl}_2}$$

b)  $V = 850. \text{ L}$

$T = 298 \text{ K}$

$P = \text{unknown}$

$n = 2.36248 \times 10^3 \text{ mol Cl}_2$

Van der Waals constants from Table 5.4:

$$a = 6.49 \frac{\text{atm}\cdot\text{L}^2}{\text{mol}^2}; \quad b = 0.0562 \frac{\text{L}}{\text{mol}}$$

$$\left( P + \frac{n^2 a}{V^2} \right) (V - nb) = nRT$$

$$P_{\text{VDW}} = \frac{nRT}{V - nb} - \frac{n^2 a}{V^2}$$

$$P_{\text{VDW}} = \frac{(2.36248 \times 10^3 \text{ mol Cl}_2) \left( 0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (298 \text{ K})}{850. \text{ L} - (2.36248 \times 10^3 \text{ mol Cl}_2) \left( 0.0562 \frac{\text{L}}{\text{mol}} \right)} - \frac{(2.36248 \times 10^3 \text{ mol Cl}_2)^2 \left( 6.49 \frac{\text{atm}\cdot\text{L}^2}{\text{mol}^2} \right)}{(850. \text{ L})^2}$$
$$= 30.4134 = \mathbf{30.4 \text{ atm}}$$

c) The engineer did not completely fill the tank. She should have filled it to  $(80.0\%/100\%)(85.0 \text{ atm}) = 68 \text{ atm}$ , but only filled it to 30.4 atm.

5.100 Plan: Use the relationship  $\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$  or  $V_2 = \frac{P_1 V_1 n_2 T_2}{P_2 n_1 T_1}$ .  $R$  is fixed.

Solution:

a) As the pressure on a fixed amount of gas ( $n$  is fixed) increases at constant temperature ( $T$  is fixed), the molecules move closer together, decreasing the volume. When the pressure is increased by a factor of 2, the volume decreases by a factor of 2 at constant temperature (Boyle's law).

$$V_2 = \frac{P_1 V_1 T_2}{P_2 T_1} = \frac{(P_1)(V_1)(1)}{(2P_1)(1)} \quad V_2 = \frac{1}{2} V_1$$

**Cylinder B** has half the volume of the original cylinder.

b) The temperature is decreased by a factor of 2, so the volume is decreased by a factor of 2 (Charles's law).

$$V_2 = \frac{P_1 V_1 T_2}{P_2 T_1} = \frac{(1)(V_1)(200 \text{ K})}{(1)(400 \text{ K})} \quad V_2 = \frac{1}{2} V_1$$

**Cylinder B** has half the volume of the original cylinder.

c)  $T_1 = 100^\circ\text{C} + 273 = 373 \text{ K}$

$T_2 = 200^\circ\text{C} + 273 = 473 \text{ K}$

The temperature increases by a factor of  $473/373 = 1.27$ , so the volume is increased by a factor of 1.27 (Charles's law).

$$V_2 = \frac{P_1 V_1 T_2}{P_2 T_1} = \frac{(1)(V_1)(473 \text{ K})}{(1)(373 \text{ K})} \quad V_2 = 1.27 V_1$$

**None** of the cylinders show a volume increase of 1.27.

d) As the number of molecules of gas increases at constant pressure and temperature ( $P$  and  $T$  are fixed), the force they exert on the container increases. This results in an increase in the volume of the container. Adding 0.1 mole

of gas to 0.1 mole increases the number of moles by a factor of 2, thus the volume increases by a factor of 2 (Avogadro's law).

$$V_2 = \frac{P_1 V_1 n_2 T_2}{P_2 n_1 T_1} = \frac{(1)(V_1)(0.2)(1)}{(1)(0.1)(1)} \quad V_2 = 2V_1$$

**Cylinder C** has a volume that is twice as great as the original cylinder.

e) Adding 0.1 mole of gas to 0.1 mole increases the number of moles by a factor of 2, thus increasing the volume by a factor of 2. Increasing the pressure by a factor of 2 results in the volume decreasing by a factor of  $\frac{1}{2}$ . The two volume changes cancel out so that the volume does not change.

$$V_2 = \frac{P_1 V_1 n_2 T_2}{P_2 n_1 T_1} = \frac{(P_1)(V_1)(0.2)(1)}{(2P_1)(0.1)(1)} \quad V_2 = V_1$$

**Cylinder D** has the same volume as the original cylinder.

- 5.101 **Plan:** Since the mole fractions of the three gases must add to 1, the mole fraction of methane is found by subtracting the sum of the mole fractions of helium and argon from 1.  $P_{\text{methane}} = X_{\text{methane}} P_{\text{total}}$  is used to calculate the pressure of methane and then the ideal gas law is used to find moles of gas. Avogadro's number is needed to convert moles of methane to molecules of methane.

**Solution:**

$$X_{\text{methane}} = 1.00 - (X_{\text{argon}} + X_{\text{helium}}) = 1.00 - (0.35 + 0.25) = 0.40$$

$$P_{\text{methane}} = X_{\text{methane}} P_{\text{total}} = (0.40)(1.75 \text{ atm}) = 0.70 \text{ atm CH}_4$$

$$V = 6.0 \text{ L}$$

$$T = 45^\circ\text{C} + 273 = 318 \text{ K}$$

$$P = 0.70 \text{ atm}$$

$$n = \text{unknown}$$

$$PV = nRT$$

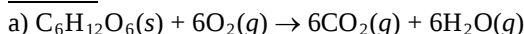
Solving for  $n$ :

$$\text{Moles of CH}_4 = n = \frac{PV}{RT} = \frac{(0.70 \text{ atm})(6.0 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(318 \text{ K})} = 0.1608715 \text{ mol}$$

$$\begin{aligned} \text{Molecules of CH}_4 &= (0.1608715 \text{ mol CH}_4) \left( \frac{6.022 \times 10^{23} \text{ CH}_4 \text{ molecules}}{1 \text{ mol CH}_4} \right) \\ &= 9.68768 \times 10^{22} = \mathbf{9.7 \times 10^{22} \text{ molecules CH}_4} \end{aligned}$$

- 5.102 **Plan:** For part a), convert mass of glucose to moles and use the molar ratio from the balanced equation to find the moles of  $\text{CO}_2$  gas produced. Use the ideal gas law to find the volume of that amount of  $\text{CO}_2$ . Pressure must be in units of atm and temperature in kelvins. For part b), use the molar ratios in the balanced equation to calculate the moles of each gas and then use Dalton's law of partial pressures to determine the pressure of each gas.

**Solution:**



$$\text{Moles CO}_2: (20.0 \text{ g C}_6\text{H}_{12}\text{O}_6) \left( \frac{1 \text{ mol C}_6\text{H}_{12}\text{O}_6}{180.16 \text{ g C}_6\text{H}_{12}\text{O}_6} \right) \left( \frac{6 \text{ mol CO}_2}{1 \text{ mol C}_6\text{H}_{12}\text{O}_6} \right) = 0.666075 \text{ mol CO}_2$$

Finding the volume of  $\text{CO}_2$ :

$$V = \text{unknown}$$

$$T = 37^\circ\text{C} + 273 = 310 \text{ K}$$

$$P = 780. \text{ torr}$$

$$n = 0.666075 \text{ mol}$$

$$\text{Converting } P \text{ from torr to atm: } P = (780. \text{ torr}) \left( \frac{1 \text{ atm}}{760 \text{ torr}} \right) = 1.0263158 \text{ atm}$$

$$PV = nRT$$

Solving for  $V$ :

$$V = \frac{nRT}{P} = \frac{(0.666075 \text{ mol}) \left( 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (310 \text{ K})}{(1.0263158 \text{ atm})} = 16.5176 = \mathbf{16.5 \text{ L CO}_2}$$

This solution assumes that partial pressure of  $\text{O}_2$  does not interfere with the reaction conditions.

$$\begin{aligned} \text{b) Moles CO}_2 &= \text{moles O}_2 = (10.0 \text{ g C}_6\text{H}_{12}\text{O}_6) \left( \frac{1 \text{ mol C}_6\text{H}_{12}\text{O}_6}{180.16 \text{ g C}_6\text{H}_{12}\text{O}_6} \right) \left( \frac{6 \text{ mol}}{1 \text{ mol C}_6\text{H}_{12}\text{O}_6} \right) \\ &= 0.333037 \text{ mol CO}_2 = \text{mol O}_2 \end{aligned}$$

At 37°C, the vapor pressure of water is 48.8 torr. No matter how much water is produced, the partial pressure of H<sub>2</sub>O will still be 48.8 torr. The remaining pressure, 780 torr – 48.8 torr = 731.2 torr is the sum of partial pressures for O<sub>2</sub> and CO<sub>2</sub>. Since the mole fractions of O<sub>2</sub> and CO<sub>2</sub> are equal, their pressures must be equal, and must be one-half of 731.2 torr.

**$P_{\text{water}} = 48.8 \text{ torr}$**

**$(731.2 \text{ torr})/2 = 365.6 = 3.7 \times 10^2 \text{ torr}$**   $P_{\text{oxygen}} = P_{\text{carbon dioxide}}$

- 5.103 **Plan:** Use the relationship between mole fraction and partial pressure,  $P_A = X_A P_{\text{total}}$ , to find the mole fraction of each gas in parts a) and b). For parts c) and d), use the ideal gas law to find the moles of air in 1000 L of air at these conditions and compare the moles of each gas to the moles of air. Mass and molecules must be converted to moles.

**Solution:**

a) Assuming the total pressure is 1 atm = 760 torr.

$$P_A = X_A P_{\text{total}}$$

$$X_{\text{Br}_2} = \frac{P_{\text{Br}_2}}{P_{\text{total}}} = \frac{0.2 \text{ torr}}{760 \text{ torr}} = 2.6315789 \times 10^{-4} \times (10^6) = 263.15789 = 300 \text{ ppmv } \textbf{Unsafe}$$

$$\text{b) } X_{\text{CO}_2} = \frac{P_{\text{CO}_2}}{P_{\text{total}}} = \frac{0.2 \text{ torr}}{760 \text{ torr}} = 2.6315789 \times 10^{-4} \times (10^6) = 263.15789 = 300 \text{ ppmv } \textbf{Safe}$$

$$(0.2 \text{ torr CO}_2 / 760 \text{ torr}) (10^6) = 263.15789 = 300 \text{ ppmv CO}_2 \textbf{ Safe}$$

$$\text{c) Moles Br}_2 = (0.0004 \text{ g Br}_2) \left( \frac{1 \text{ mol Br}_2}{159.80 \text{ g Br}_2} \right) = 2.5031 \times 10^{-6} \text{ mol Br}_2 \text{ (unrounded)}$$

Finding the moles of air:

$$V = 1000 \text{ L}$$

$$P = 1.00 \text{ atm}$$

$$PV = nRT$$

$$T = 0^\circ\text{C} + 273 = 273 \text{ K}$$

$$n = \text{unknown}$$

$$\text{Moles of air} = n = \frac{PV}{RT} = \frac{(1.00 \text{ atm})(1000 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(273 \text{ K})} = 44.616 \text{ mol air (unrounded)}$$

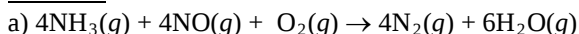
$$\begin{aligned} \text{Concentration of Br}_2 &= \text{mol Br}_2 / \text{mol air} (10^6) = [(2.5031 \times 10^{-6} \text{ mol}) / (44.616 \text{ mol})] (10^6) \\ &= 0.056103 = 0.06 \text{ ppmv Br}_2 \textbf{ Safe} \end{aligned}$$

$$\text{d) Moles CO}_2 = (2.8 \times 10^{22} \text{ molecules CO}_2) \left( \frac{1 \text{ mol CO}_2}{6.022 \times 10^{23} \text{ molecules CO}_2} \right) = 0.046496 \text{ mol CO}_2$$

$$\begin{aligned} \text{Concentration of CO}_2 &= \text{mol CO}_2 / \text{mol air} (10^6) = [(0.046496 \text{ mol}) / (44.616 \text{ mol})] (10^6) = 1042.1 \\ &= 1.0 \times 10^3 \text{ ppmv CO}_2 \textbf{ Safe} \end{aligned}$$

- 5.104 **Plan:** For part a), use the ideal gas law to find the moles of NO in the flue gas. The moles of NO are converted to moles of NH<sub>3</sub> using the molar ratio in the balanced equation and the moles of NH<sub>3</sub> are converted to volume using the ideal gas law. For part b), the moles of NO in 1 kL of flue gas is found using the ideal gas law; the molar ratio in the balanced equation is used to convert moles of NO to moles and then mass of NH<sub>3</sub>.

**Solution:**



Finding the moles of NO in 1.00 L of flue gas:

$$V = 1.00 \text{ L}$$

$$P = 4.5 \times 10^{-5} \text{ atm}$$

$$PV = nRT$$

Solving for  $n$ :

$$T = 365^\circ\text{C} + 273 = 638 \text{ K}$$

$$n = \text{unknown}$$

$$\text{Moles of NO} = n = \frac{PV}{RT} = \frac{(4.5 \times 10^{-5} \text{ atm})(1.00 \text{ L})}{\left(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(638 \text{ K})} = 8.5911 \times 10^{-7} \text{ mol NO}$$

$$\text{Moles of NH}_3 = (8.5911 \times 10^{-7} \text{ mol NO}) \left( \frac{4 \text{ mol NH}_3}{4 \text{ mol NO}} \right) = 8.5911 \times 10^{-7} \text{ mol NH}_3$$

Volume of NH<sub>3</sub>:

V = unknown

$$T = 365^\circ\text{C} + 273 = 638 \text{ K}$$

P = 1.00 atm

$$n = 8.5911 \times 10^{-7} \text{ mol}$$

PV = nRT

Solving for V:

$$V = \frac{nRT}{P} = \frac{(8.5911 \times 10^{-7} \text{ mol}) \left(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(638 \text{ K})}{(1.00 \text{ atm})} = 4.5000 \times 10^{-5} = \mathbf{4.5 \times 10^{-5} \text{ L NH}_3}$$

b) Finding the moles of NO in 1.00 kL of flue gas:

V = 1.00 kL = 1000 L

$$T = 365^\circ\text{C} + 273 = 638 \text{ K}$$

P = 4.5 × 10<sup>-5</sup> atm

n = unknown

PV = nRT

Solving for n:

$$\text{Moles of NO} = n = \frac{PV}{RT} = \frac{(4.5 \times 10^{-5} \text{ atm})(1000 \text{ L})}{\left(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(638 \text{ K})} = 8.5911 \times 10^{-4} \text{ mol NO}$$

$$\text{Moles of NH}_3 = (8.5911 \times 10^{-4} \text{ mol NO}) \left( \frac{4 \text{ mol NH}_3}{4 \text{ mol NO}} \right) = 8.5911 \times 10^{-4} \text{ mol NH}_3$$

$$\text{Mass of NH}_3 = (8.59 \times 10^{-4} \text{ mol NH}_3) \left( \frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3} \right) = 0.014631 = \mathbf{0.015 \text{ g NH}_3}$$

5.105 Plan: Use Graham's law to compare effusion rates.

Solution:

$$\frac{\text{Rate Ne}}{\text{Rate Xe}} = \sqrt{\frac{\text{molar mass Xe}}{\text{molar mass Ne}}} = \sqrt{\frac{131.3 \text{ g/mol}}{20.18 \text{ g/mol}}} = \frac{2.55077}{1} \text{ enrichment factor (unrounded)}$$

$$\text{Thus } X_{\text{Ne}} = \frac{\text{moles of Ne}}{\text{moles of Ne} + \text{moles of Xe}} = \frac{2.55077 \text{ mol}}{2.55077 \text{ mol} + 1 \text{ mol}} = 0.71837 = \mathbf{0.7184}$$

5.106 Plan: To find the number of steps through the membrane, calculate the molar masses to find the ratio of effusion rates. This ratio is the enrichment factor for each step.

Solution:

$$\begin{aligned} \frac{\text{Rate}_{^{235}\text{UF}_6}}{\text{Rate}_{^{238}\text{UF}_6}} &= \sqrt{\frac{\text{molar mass } ^{238}\text{UF}_6}{\text{molar mass } ^{235}\text{UF}_6}} = \sqrt{\frac{352.04 \text{ g/mol}}{349.03 \text{ g/mol}}} \\ &= 1.004302694 \text{ enrichment factor} \end{aligned}$$

Therefore, the abundance of <sup>235</sup>UF<sub>6</sub> after one membrane is 0.72% × 1.004302694

Abundance of <sup>235</sup>UF<sub>6</sub> after "N" membranes = 0.72% × (1.004302694)<sup>N</sup>

Desired abundance of <sup>235</sup>UF<sub>6</sub> = 3.0% = 0.72% × (1.004302694)<sup>N</sup>

Solving for N:

$$3.0\% = 0.72\% \times (1.004302694)^N$$

$$4.16667 = (1.004302694)^N$$

$$\ln 4.16667 = \ln (1.004302694)^N$$

$$\ln 4.16667 = N \times \ln (1.004302694)$$

$$N = (\ln 4.16667)/(\ln 1.004302694)$$

$$N = 1.4271164/0.004293464 = 332.39277 = \mathbf{332 \text{ steps}}$$

- 5.107 Plan: The amount of each gas that leaks from the balloon is proportional to its effusion rate. Using 35% as the rate for H<sub>2</sub>, the rate for O<sub>2</sub> can be determined from Graham's law.

Solution:

$$\frac{\text{Rate O}_2}{\text{Rate H}_2} = \sqrt{\frac{\text{molar mass H}_2}{\text{molar mass O}_2}} = \sqrt{\frac{2.016 \text{ g/mol}}{32.00 \text{ g/mol}}} = \frac{\text{rate O}_2}{35}$$

$$0.250998008 = \frac{\text{rate O}_2}{35}$$

$$\text{Rate O}_2 = 8.78493$$

Amount of H<sub>2</sub> that leaks = 35%; 100–35 = 65% H<sub>2</sub> remains

Amount of O<sub>2</sub> that leaks = 8.78493%; 100–8.78493 = 91.21507% O<sub>2</sub> remains

$$\frac{\text{O}_2}{\text{H}_2} = \frac{91.21507}{65} = 1.40331 = \mathbf{1.4}$$

- 5.108 Plan: For part a), put together the various combinations of the two isotopes of Cl with P and add the masses. Multiply the abundances of the isotopes in each combination to find the most abundant for part b). For part c), use Graham's law to find the effusion rates.

Solution:

a) Options for PCl<sub>3</sub>:

All values are g/mol

| P  | First Cl | Second Cl | Third Cl | Total |
|----|----------|-----------|----------|-------|
| 31 | 35       | 35        | 35       | 136   |
| 31 | 37       | 35        | 35       | 138   |
| 31 | 37       | 37        | 35       | 140   |
| 31 | 37       | 37        | 37       | 142   |

b) The fraction abundances are <sup>35</sup>Cl = 75%/100% = 0.75, and <sup>37</sup>Cl = 25%/100% = 0.25.

The relative amount of each mass comes from the product of the relative abundances of each Cl isotope.

$$\mathbf{\text{Mass 136} = (0.75) (0.75) (0.75) = 0.421875 = 0.42 \text{ (most abundant)}}$$

$$\text{Mass 138} = (0.25) (0.75) (0.75) = 0.140625 = 0.14$$

$$\text{Mass 140} = (0.25) (0.25) (0.75) = 0.046875 = 0.047$$

$$\text{Mass 142} = (0.25) (0.25) (0.25) = 0.015625 = 0.016$$

$$\text{c) } \frac{\text{Rate P}^{37}\text{Cl}_3}{\text{Rate P}^{35}\text{Cl}_3} = \sqrt{\frac{\text{molar mass P}^{35}\text{Cl}_3}{\text{molar mass P}^{37}\text{Cl}_3}} = \sqrt{\frac{136 \text{ g/mol}}{142 \text{ g/mol}}}$$

$$= 0.978645 = \mathbf{0.979}$$